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ABSTRACT

Designed to accompany the student text on respiration, this manual presents instructions on the use of laboratory equipment and presents various experiments dealing with the concepts presented in the text. Thirty-nine laboratory activities are described. Laboratory activities are divided into several parts, each part covering a specific experiment dealing with the concept covered by the activity. Each experiment includes a description of materials, procedures, and discussion questions. (RE)

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# BIOMEDICAL SCIENCE

## UNIT I

# RESPIRATION IN HEALTH AND MEDICINE

## RESPIRATORY ANATOMY, PHYSIOLOGY AND PATHOLOGY; THE BEHAVIOR OF GASES; INTRODUCTORY CHEMISTRY; AND AIR POLLUTION

LABORATORY MANUAL
REVISED VERSION, 1975

THE BIOMEDICAL INTERDISCIPLINARY CURRICULUM PROJECT

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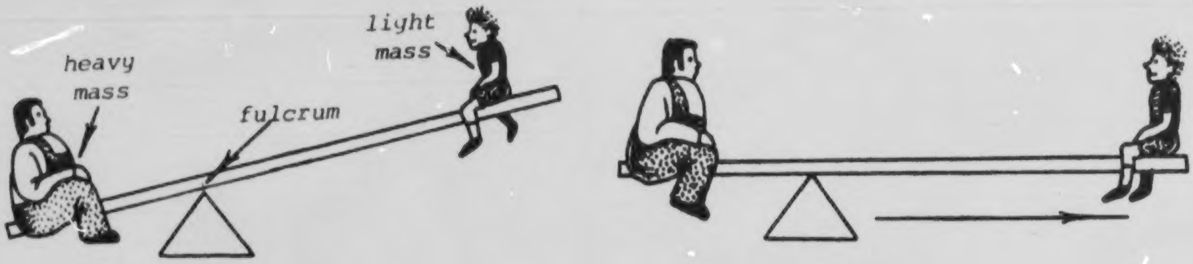
## LABORATORY ACTIVITY 2:

## CARE AND FEEDING OF THE LABORATORY BALANCE

### GENERAL INTRODUCTION:

Laboratory balances come in a variety of types and designs, but they all operate on the same principle. They all are designed to balance an unknown mass against a known mass. The underlying principle is the same one that children use when playing on a see-saw. Every child learns that a lighter child can balance a heavier one by sitting farther away from the fulcrum.



*Unequal masses out of balance, both the same distance from the fulcrum.*

*Lighter mass has been shifted to a position that gives a balance.*

FIGURE 1: *Balancing unequal masses.*

Notice that if the lighter mass in Figure 1 were shifted even more to the right, the platform would go out of balance in the opposite direction. In other words, if the position of one mass is fixed there is only one specific position of the second mass that will provide a balance between the two masses.

If the distance of each mass from the fulcrum is known and we also know the value of one mass, the value of the second mass can be calculated. None of this is necessary, however, in using a laboratory balance.

With a laboratory balance, the unknown mass is placed in a pan (or on a platform) that has a fixed position in relation to the fulcrum. On the other side of the fulcrum are a set of movable masses (usually three of them). The position of these masses is adjusted until they balance the unknown mass. Once a balance is achieved, the value of the unknown mass can be read directly from the instrument.

The purpose of this activity is to acquaint you with the nature of the laboratory balance and the techniques involved in using it to determine the mass of different kinds of substances. The procedure is divided into four short parts.

## PART I: THE NATURE OF THE BALANCE

INTRODUCTION:

It is important to understand that the laboratory balance is an expensive and delicate instrument that is capable of providing valuable information. As such, it should be treated carefully and with respect. The balance should be kept clean at all times. Since many chemicals can corrode the balance, spills that are not cleaned up promptly can decrease the accuracy of the instrument.

Before actually weighing an object, it will be necessary to be able to recognize the parts of the balance. The purpose of Part I is to learn about the components of the balance.

MATERIALS:

balance

PROCEDURE:

1. Although there are many different types of balances, they all have certain features in common. Compare the balance you will be using with Figure 2. Locate the labeled parts on your balance. The balance should have a weighing pan that is connected to horizontal beams. On the beams are sliding riders. In laboratory use, materials placed in the weighing pan are counterbalanced by the riders that are moved to different positions along the beams. The total mass in grams is determined by adding up the numbers marked by the riders. The largest amount that can be weighed is the sum of the highest numbers on the beams. What is the maximum capacity in grams of the balance you will be using?

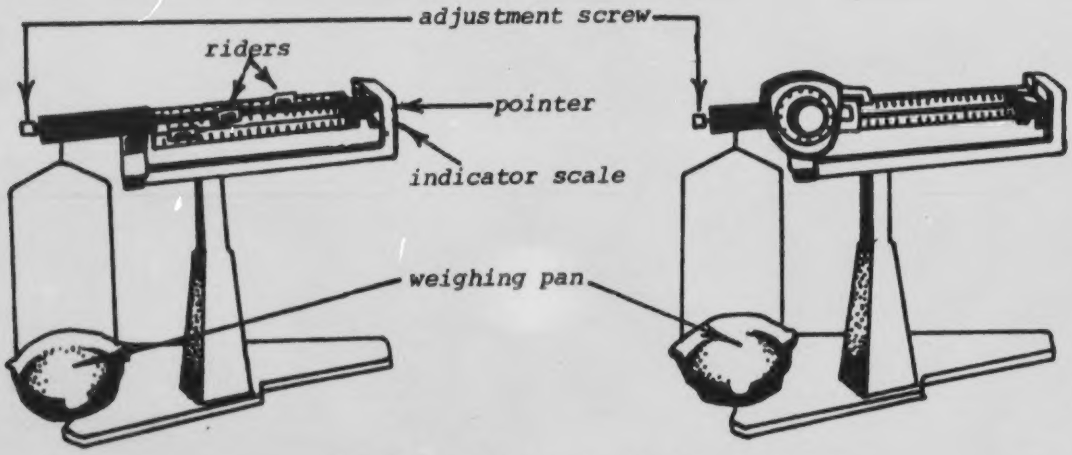


FIGURE 2: *Parts of a balance.*

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2. Before any measurements can be made, the balance should be adjusted so that it measures zero when the weighing pan is empty. Move all of the riders to the "zero mark" and empty the weighing pan. The pointer should swing back and forth slowly across the middle of the indicator scale. It will not be necessary to wait for the pointer to stop completely. A swing of an equal number of divisions on either side of the zero mark on the indicator scale indicates that the balance is properly leveled. If the pointer does not swing properly, the adjustment screw should be turned carefully until the balance is correctly adjusted.

## PART II: THE MASS OF A BEAKER

### INTRODUCTION:

In this part of the activity, a 50-ml beaker will be weighed several times. This is the first of three weighing tasks you will be asked to perform. The object will be to gain some experience in the use of the balance, for this instrument will be used in many activities in the Biomedical Science course.

### MATERIALS:

balance
beaker, 50-ml

### PROCEDURE:

1. Place a dry 50-ml beaker on the weighing pan. The beaker can be weighed by moving the riders until the pointer rests near the middle of the indicator scale (Figure 2). Remember that you do not need to wait for the pointer to come to a complete rest (see Part I, Step 2).

2. With all riders at the zero mark and with the beaker on the pan, the pointer should be above the middle of the scale. Start with the heaviest rider and move it from the zero mark to the first notch on the beam. If the pointer remains above the middle of the scale, that means that the object on the pan weighs more than the number near the notch. If so, move the same rider to the next notch, and continue moving the rider from notch to notch until the pointer falls below the middle of the scale. That tells you the range of the mass of the beaker as in the following examples.

### EXAMPLE 1:

If the pointer moves below the middle of the scale when the rider is transferred from the zero mark to the 10 g notch, then the beaker weighs less than 10 g.

EXAMPLE 2:

If the pointer moves below the middle of the scale when the rider is moved from the 20 g to the 30 g notch, then the beaker weighs between 20 and 30 g.

3. Move the rider back to the notch position just before the pointer fell below the middle of the scale. (In example 2, the rider would be moved back to the 20-g notch.) Adjust the position of the middle rider from notch to notch just as you did in Step 2--until the pointer again falls below the middle of the scale. Then move the rider back one notch.

4. Finally, adjust the position of the lightest rider until the pointer is balanced--swings equally on either side of the middle of the indicator scale.

5. The total of the masses indicated by the riders is the mass of the beaker. Record this mass to the nearest .01 g on a data sheet. To show you a convenient way to record data for this activity, a sample data sheet is provided at the end of the activity. Copy the data sheet on a separate sheet of paper. Your name should be on the sheet because your instructor may wish to review your results.

6. Return all riders to the zero mark and reweigh the same beaker at least six times (i.e., repeat Steps 1 through 5). In each case, try to ignore your earlier results--they may be inaccurate. Be sure to record the mass of the beaker in each weighing. These data will be used in a later class period. Use your own data sheet for keeping records.

7. To get a feel for the sensitivity of the balance, determine how small a movement of the lightest rider can be detected by a change in the position of the pointer. Record this information on your own data sheet.

## PART III: THE MASS OF A SOLID

### INTRODUCTION:

The laboratory balance is frequently used to measure quantities of powdered or granular solids. In this part of the activity, you will weigh out some sand.

It is not recommended that chemicals be placed directly on the weighing pan. Instead, the chemicals are placed on a piece of paper (a tare) on the weighing pan in order to keep the balance clean.

Another tool useful for measuring solids is the scoopula. Laboratory chemicals should never be handled directly with your fingers, since some of them are dangerous and can cause injury. A scoopula is therefore used to transfer chemicals from a container to the tare. If scoopulas are not available, plastic spoons or laboratory spatulas serve as acceptable substitutes. Finally, to avoid contamination of the chemical, it is essential to keep the scoopulas as clean as possible.

MATERIALS:

balance
tare
scoopula (or plastic spoon)
sand
beaker, 50-ml

PROCEDURE:

1. Obtain a tare and place it on the weighing pan. Observe what happens to the pointer when the riders are set at zero. Notice that the tare weighs a certain amount that must be accounted for. For example, if 3.0 g of a certain chemical are needed, then the riders should be set at a number equal to 3.0 g more than the mass of the tare. If the tare weighs 0.3 g, then a 3.0 g quantity of a solid can be weighed by setting the riders so that they add up to 3.3 g. The pointer will stop at the middle of the indicator scale when 3.0 g of the chemical have been transferred to the tare.

2. Fill a 50-ml beaker approximately half full with sand. With the help of the following directions, weigh out 5.00 g of sand.

a. First measure the mass of a tare to the nearest 0.01 g and record this amount on your own data sheet. A sample data sheet is shown at the end of the activity. Add this amount to 5.00 g so that the riders can be adjusted to the proper position.

b. Using a scoopula, transfer the sand to the tare until the instrument is balanced. If too much is added, the pointer will swing above the middle of the indicator scale. When this happens, simply remove small amounts of sand with the scoopula until the balance is properly adjusted. The tare should now contain 5.00 grams of sand.

3. Follow your instructor's directions on disposing of the sand.

## PART IV: MASS OF A LIQUID

INTRODUCTION:

The balance is also frequently used to measure the mass of liquids. Since a paper tare cannot hold liquids, beakers or other containers are used as tares.

MATERIALS:

balance
2 beakers, 50-ml
medicine dropper

PROCEDURE:

1. Reweigh the beaker used in Part III. It will be used as a tare for weighing out 20.00 g of water. The beaker should be clean and dry at this point.

2. Adjust the riders until the beaker is balanced and record the mass on your own data sheet. A sample sheet is shown at the end of the activity.

3. Using another beaker, try to add 20.00 g of water to the beaker on the weighing pan. Remember to adjust the riders (see sample data sheet, Part IV) before adding the water.

4. This is actually a very challenging task. You may have to settle for a mass of water that is slightly different from 20.00 g. Small amounts of water may be added or removed with a medicine dropper. When you have come as close to 20.00 g as possible, find the mass of the water in the beaker by adjusting the lightest rider. Record the mass and then subtract the mass of the beaker.

5. Now pour out all of the water and reweigh the same beaker. Record the mass of the wet beaker and compare it with the previous "dry mass." The change in mass is due to the small amount of water that remains on the sides of the container. When highly precise measurements are needed, this factor should always be kept in mind.

6. Find and record the mass of one drop of water.

SAMPLE DATA SHEET:

PART II: MASS OF A BEAKER

| Trial Number | Mass of Beaker in grams |
|---|---|
| 1 | |
| 2 | |

PART III: MASS OF A SOLID

Mass of tare . . . . . . . . . _____ grams
Desired mass of sand . . . . + 5.00 grams
Riders should be set to. . . grams total

## PART IV: MASS OF A LIQUID

Mass of beaker (tare). . . . _____ grams

Desired mass of water. . . . + 20.00 grams

Riders should be set to. . . grams total

Actual mass of water
and beaker . . . . . . . . _____ grams

Mass of beaker . . . . . . . . - _____ grams

Mass of water obtained . . . _____ grams

Mass of wet beaker . . . . . _____ grams

Mass of one drop of water. . _____ grams

## LABORATORY ACTIVITY 3:

## VOLUMETRIC MEASUREMENTS

### GENERAL INTRODUCTION:

This activity studies two common instruments used for measuring volumes of liquids, the graduated cylinder (or graduate) and the graduated pipet. Correct measuring technique is essential in all laboratory work both in the classroom and in clinical laboratories. This activity is divided into three parts. In the first two sections, you will study the graduated cylinder and the pipet, respectively, and gain some familiarity with their use. In the last section, the precision of the pipet will be compared to that of a graduated cylinder.

### PART I: USE OF THE GRADUATED CYLINDER

### MATERIALS:

| | |
|---|---|
| graduated cylinder, 100-ml | 2 beakers, 50-ml |
| medicine dropper | balance |

### PROCEDURE:

1. Examine the 100-ml graduated cylinder. It is designed to contain one tenth of a liter, and is divided into 100 equal parts (ml). Using a beaker, fill the graduate with water to the 25-ml mark. If you examine it closely, you will notice that the surface of the water is curved (see Figure 1). This curve is called the meniscus. The graduate is designed so that it contains 25 ml when the bottom of the meniscus is even with the 25-ml line (again see Figure 1). Check to see whether your graduate contains 25 ml. If necessary, additional water

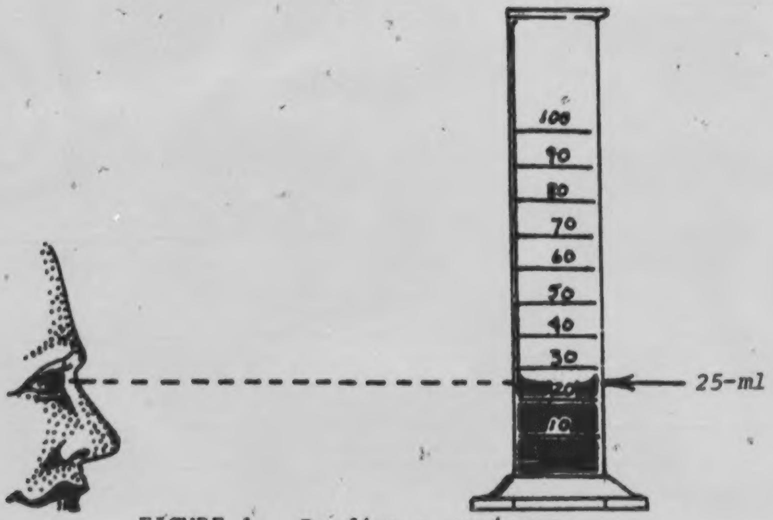


FIGURE 1: *Reading a meniscus.*

may be added, drop-by-drop, with a medicine dropper. Note: it is important to keep your eyes at the level of the meniscus. The accuracy will be reduced if the graduate is filled while viewing from above or below this level. It is sometimes helpful to hold a dark object behind the graduate such as a black card in order to see the meniscus clearly.

2. Using a clean, dry beaker, weigh the 25 ml of water. First weigh the beaker to the nearest 0.01 g and record its mass on a data sheet as you did in the previous lab activity. (A sample data sheet is shown at the end of the activity.) Then, carefully pour the water from the graduate into the beaker and adjust the riders so that the balance is leveled. Determine and record the mass of the water by subtracting the mass of the beaker from the combined mass of the beaker and the water. If you're not sure how to do this, review Laboratory Activity 2.

DISCUSSION QUESTIONS:

1. What relationship do you suppose might exist between milliliters of water and grams of water? (Hint: Remember that it is not possible to pour all the water out of a container.)

2. Can you suggest an alternate method for measuring volumes of water?

## PART II: USE OF THE GRADUATED PIPET

INTRODUCTION:

Pipets are normally used to measure volumes that are less than 10 ml. Compare the two different sizes of pipets. One holds 10 ml and has 0.1-ml divisions, while the other holds 1 ml and has 0.01-ml divisions. Pipets are generally filled using mouth suction in much the same way that a drinking straw is used. However, laboratory chemicals should never be drawn into the mouth since they are sometimes very harmful. You will not be asked to pipet anything by mouth that is dangerous. If there is a need to transfer a poisonous or dangerous fluid, other techniques will be introduced.

Notice that the pipet is graduated with the numbers running from top to bottom. If you desire to measure 8.0 ml of water, the 10-ml pipet should be filled to the line marked "2." The liquid is then allowed to flow out of the pipet into an appropriate receiving vessel, such as a beaker. A small amount of water will remain in the tip of the pipet, which should always be blown out into the beaker to complete the transfer.

The following is a general guide for pipetting (see Figure 2). It will be helpful to read the entire procedure before beginning. Practice pipetting different volumes of water until you understand the technique thoroughly.

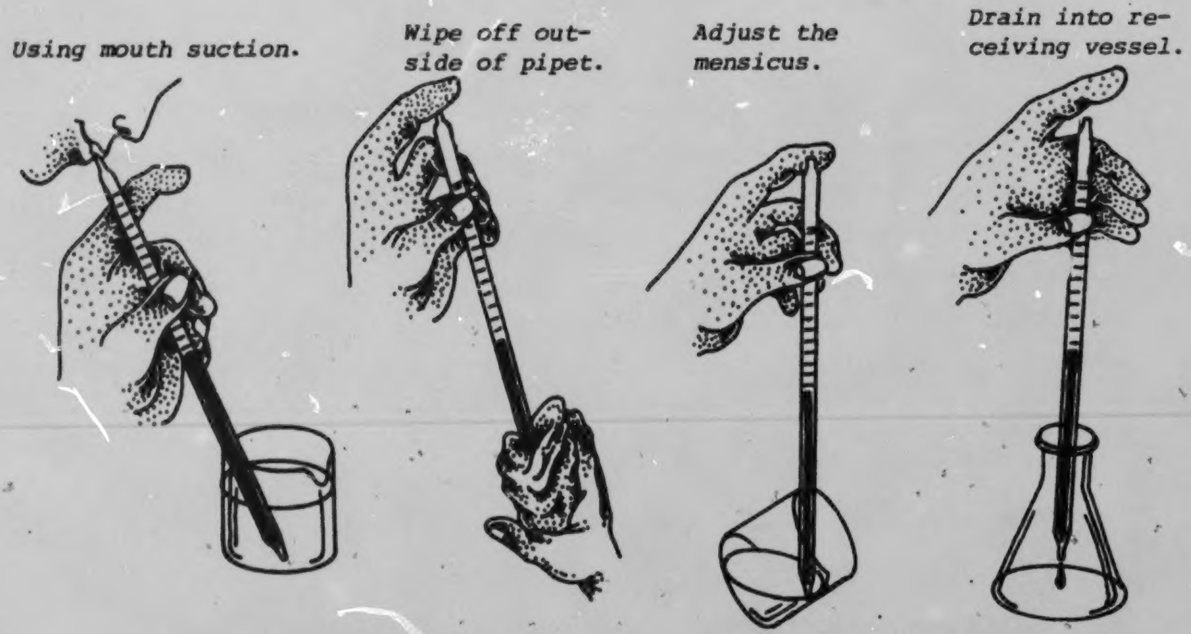

Using mouth suction.
Wipe off outside of pipet.
Adjust the mensicus.
Drain into receiving vessel.

FIGURE 2: *Pipetting technique.*

PROCEDURE:

1. Check the pipet to make sure that it is the correct size for the volume that must be measured. Decide to what level the pipet should be filled before beginning.

2. Hold the pipet with the thumb and the last three fingers, keeping the index finger free as in Figure 2.

3. Place (and keep) the tip of the pipet below the surface of the liquid to be transferred, and, using mouth suction, slowly draw the liquid up to at least one inch above the desired level.

4. Remove your mouth and quickly place the index finger over the end of the pipet. This may take some practice to succeed. The index finger should prevent the liquid from coming out.

5. Briefly wipe off the tip of the pipet with a tissue. A drop of fluid on the outside of the pipet could lead to a serious error in some activities.

6. With the tip of the pipet touching the wall of the original container (not the liquid), raise your index finger slightly so that the meniscus drops

slowly to the desired level. (Using a slight rolling motion of the fingertip may help.) Remember to keep the eye at the same level as the meniscus, and to hold the pipet vertically while making the measurement.

7. Place the pipet in the new container, and allow the water to drain out by removing your finger. Gently blow the small amount remaining in the tip of the pipet into the new container. The transfer is now complete. The entire process may seem a little awkward at first but with care and practice, you can master the technique in a short time.

An alternate method for measuring small volumes is to allow the liquid to drain from one level to another, without letting all of the fluid drain out. For example, 4.0 ml of water might be transferred by letting the meniscus drop from zero to four or from two to six. This technique should also be practiced since it is quite useful.

Certain general considerations should always be kept in mind while pipetting. Do not pipet fluids into your mouth. But if a solution (other than water) accidentally is drawn into the mouth, it should be immediately spit out and the mouth rinsed several times with water.

## PART III: COMPARING THE PRECISION OF THE GRADUATED CYLINDER AND THE PIPET

### MATERIALS:

graduated cylinder, 10-ml
pipet, 10-ml
2 beakers, 50-ml
medicine dropper
balance

### PROCEDURE:

1. Weigh a dry beaker to the nearest 0.01 g with the balance, and record the mass on your own data sheet.

2. With the 10-ml graduate, measure 5.0 ml of water and pour it into the beaker. This will be called Sample I. Weigh the water and the beaker and record their total mass. A useful form for handling the data of Part III is shown in the sample data sheet at the end of the activity.

3. Using the 10-ml graduate, transfer 5.0 more ml of water into the same beaker. The beaker should now contain 10.0 ml of water. Record the new total mass.

4. Repeat Step 3 so that the beaker contains 15.0 ml of water. Record the new total mass.

5. Repeat Steps 2 through 4, except that this time use a 10-ml pipet instead of a graduated cylinder. Record the data. These are numbered 4 to 6 on the sample data sheet.

6. Perform the subtractions indicated for the six samples. By comparing the results obtained using the graduate with those obtained with the pipet, it should be possible to make some preliminary observations concerning the precision of each.

DISCUSSION QUESTIONS:

1. Why was the water left in the beaker during the procedure steps of Part III?

2. Which instrument do you feel is better for measuring volumes of liquids? Why?

3. Why is the tip of the pipet wiped off before adjusting the meniscus?

SAMPLE DATA SHEET:

### PART I: GRADUATED CYLINDER MEASUREMENT

MASS OF:
beaker + 25.0 ml water. . ________ g
beaker. . . . . . . . . . -________ g
25 ml water . . . . . . . ________ g

### PART III: GRADUATED CYLINDER MEASUREMENTS

MASS OF:
beaker + 5.0 ml water . . ________ g
beaker. . . . . . . . . . -________ g
SAMPLE 1. . . . . . . . . ________ g

MASS OF:
beaker + 10.0 ml water. . ________ g
beaker + 5.0 ml water . . -________ g
SAMPLE 2. . . . . . . . . ________ g

MASS OF:
beaker + 15.0 ml water. . ________ g
beaker + 10.0 ml water. . -________ g
SAMPLE 3. . . . . . . . . ________ g

## PART III: PIPET MEASUREMENTS

MASS OF:
beaker + 20.0 ml water . . ________ g
beaker + 15.0 ml water . . -________ g
SAMPLE 4 . . . . . . . . . . ________ g

MASS OF:
beaker + 25.0 ml water . . ________ g
beaker + 20.0 ml water . . -________ g
SAMPLE 5 . . . . . . . . . . ________ g

MASS OF:
beaker + 30.0 ml water . . ________ g
beaker + 25.0 ml water . . -________ g
SAMPLE 6 . . . . . . . . . . ________ g

## LABORATORY ACTIVITY 4:

## DENSITY

### INTRODUCTION:

The density of a substance is the amount of its mass that occupies a given amount of space. Since space is measured using volume units, density may be expressed in units of grams per milliliter. As you have seen in Laboratory Activity 3, the mass of 1 ml of water is very near 1 gram--the density of water is therefore approximately 1 g per ml.

In this activity, the densities of several liquids and solids will be measured using some of the techniques introduced in the two preceding activities. The graduated cylinder and pipet will be used to measure volumes, while the balance will be used for determining mass. The density of each substance will be calculated by dividing its mass by its volume. For example, if the mass of 10 ml of a substance is 136 grams, its density is 136 ÷ 10 = 13.6 g/ml.

### PART I: DENSITIES OF VARIOUS SOLIDS

### MATERIALS:

graduated cylinder, 100-ml
balance
2 or 3 rubber stoppers
2 or 3 glass rods
2 or 3 iron nails

### PROCEDURE:

1. Obtain two or three rubber stoppers. If there are not enough rubber stoppers for everyone, you may wish to begin with the glass rods or the iron nails. It makes no difference. Determine, to the nearest 0.01 g, the mass of the stoppers, and record this amount on your data sheet. You may wish to refer to the sample data sheet at the end of the activity.

2. Fill a 100-ml graduate to the 50-ml mark with water. Remember to use the bottom of the meniscus and to keep all readings at eye level.

3. Place the rubber stoppers in the graduate so that they are completely immersed by the water. The water level will rise. The volume of the stoppers is equal to the increase in volume indicated by the rise of the meniscus. Record the final volume in the data sheet.

4. Empty the graduate and return the stoppers to the stock table. They should be blotted dry with a paper towel before they are returned.

5. Repeat Steps 1 to 4, except make the measurements on a different substance (iron nails or glass). Remember to immerse the items completely in the water. The ends of the glass tubing should not extend above the surface. Also, it will increase the accuracy of your measurements if the items are dry when they are weighed.

6. With the data collected in Steps 1 to 5, calculate the density of the substances. If time permits, the density of other available items may be measured.

DISCUSSION QUESTIONS:

1. If you were to design an anchor for a boat, and you were asked to construct it out of one of the solids studied in Part I, which material would you use? Explain. (Assume cost is not a concern.)

2. Determination of density requires a mass measurement and a volume measurement. Which of these two measurements is likely to be less precise?

## PART II: MEASURING THE DENSITY OF VARIOUS LIQUIDS

MATERIALS:

| | |
|---|---|
| 2 pipets, 10-ml | alcohol |
| balance | cooking oil |
| 2 beakers, 50-ml | corn syrup |

PROCEDURE:

1. Measure the mass of a 50-ml beaker to the nearest 0.01 g and record this amount in the data sheet.

2. Obtain from the stock table a sample of alcohol, cooking oil or corn syrup. Pipet 10.0 ml of this liquid into the beaker.

3. Determine the mass, to the nearest 0.01 g, of the beaker plus 10.0 ml of liquid. Record this amount in the data sheet.

4. Repeat Steps 1 to 3 for the other liquids. As you learned in the previous activity, it is not necessary to empty the beaker used for weighing one substance before weighing another since you are only concerned with the increase in mass. Just remember to keep you figures straight. However, a clean pipet should be used each time.

5. Calculate the mass of each liquid using the data collected. Calculate the density of each liquid by dividing its mass by the space it occupies, which is 10 ml.

6. Wash out the pipets and beakers in order to remove all traces of cooking oil, etc.

DISCUSSION QUESTIONS:

1. Which was the most dense substance tested? Which was the least dense?

2. Explain why it was not necessary to wash out and dry the beaker between measurements of the masses of the three different liquids?

3. When cooking oil is added to water, two layers are formed. One layer is water, the other is cooking oil. Which layer would you expect to find on the bottom and which on the top? Why?

SAMPLE DATA SHEET:

PART I: DENSITY OF SOLIDS

| | Rubber |
|---|---|
| Mass of substance (g) | |
| Final volume of liquid (ml) | |
| Initial volume of liquid (ml) | 50 ml |
| Volume of substance (Final vol. - initial vol.) (ml) | |
| Density of substance (g/ml) | |

PART II: DENSITY OF LIQUIDS

| | Alcohol |
|---|---|
| Mass of beaker + 10 ml of liquid (g) | |
| Mass of beaker (g) | |
| Mass of 10 ml of liquid (g) | |
| Density of liquid (g/ml) | |

## LABORATORY ACTIVITY 5:

## NEW WAYS TO DETERMINE VOLUME

### INTRODUCTION:

In Laboratory Activity 4, the volumes of certain solids were determined by immersing them in water contained in a graduated cylinder. The volume was obtained by subtracting the original water level from the water level after submersion. The mass and the volume were then used to calculate the density of the solid in g/ml.

Finding the volume of an object by submerging it is feasible only if the object is reasonably small. For example, it would be impossible to find the volume of a room in this way. However, the volume may easily be found with the help of a meter stick and a few calculations.

This activity poses three separate problems. In the first part, you will be asked to determine the mass, surface area, volume and density of a rectangular solid. In the second part, you will determine the surface area and volume of a graduated cylinder from two measurements of length. In the last part, you are to find a way to determine the volume of an irregular solid that is too large to fit inside of a graduated cylinder.

### MATERIALS:

graduated cylinder, 100-ml
balance
millimeter ruler
rectangular solid
irregular solid

### PROCEDURE:

1. Use a millimeter ruler to measure the length, width and height of a rectangular solid. Record each value to the nearest 0.1 cm.

2. Determine and record the mass of the rectangular solid to the nearest 0.1 g.

3. Calculate the total suface area of the rectangular solid. The total surface area is the sum of the areas of the six faces of the solid. Be sure to include units in your result.

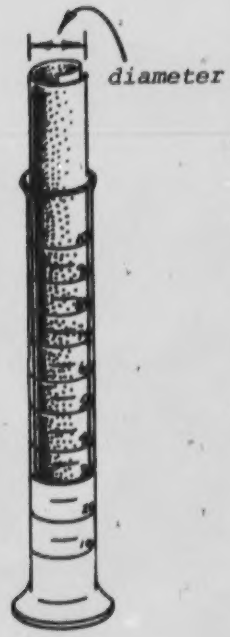
4. Use the following formula to calculate the volume of the rectangular solid. Again, include units in your result.

volume = length x width x height

5. Calculate the density of the rectangular solid.

6. State the specific gravity of the rectangular solid.

7. If we *carefully fill a graduated cylinder* to the 100-ml mark, we know that the volume of liquid will be very close to 100 ml. It is also possible to calculate the volume of liquid inside the cylinder from two length measurements. The volume of a cylinder (or a cylindrical column of liquid) is given by the formula $V = \pi r^2 h$, where $\pi = 3.14$, r is the radius and h is the height of the cylinder. Since diameter = 2 x radius, the radius of the cylinder can be found by measuring its diameter and taking half of the result.



8. Because the lip of the cylinder curves outward slightly, it is not possible to determine the diameter simply by placing a ruler across the top of the cylinder. This problem can be solved with a piece of paper rolled up inside of the cylinder, as shown in the diagram. If the paper is kept straight and close to the walls of the cylinder, a reasonably precise measurement of the diameter can be made. Use this technique to measure and record the diameter of the cylinder to the nearest 0.05 cm.

9. Measure the height of the cylinder from the "zero" position to the 100-ml mark. Record the height to the nearest 0.1 cm.

10. Calculate and record the radius of the cylinder.

11. Calculate and record the volume of the cylinder.

12. If the graduated cylinder were holding 100 ml of water, what would be the total surface area of the water? The area of the vertical walls of the column of water would be the circumference times the height. Since circumference = $2\pi r$, this area may be stated as $2\pi rh$. To this must be added the two circular areas on the bottom and the top of the column of water. Each of these circles has an area of $\pi r^2$. So the total surface area of the water is given by the formula

$$A = 2\pi rh + 2\pi r^2$$
$$= 2\pi r(h + r)$$

13. Calculate and record the total surface area of 100 ml of water in a 100-ml graduated cylinder.

14. Use your ingenuity to determine the volume of an irregular solid. Record all measurements and calculations used in obtaining the volume. (Hint: If you plan to both weigh and immerse the solid, it is important to weigh before wetting.)

## DISCUSSION QUESTIONS:

1. Describe three different methods that can be used to determine the volume of at least some kinds of solids.

2. Of the three methods, which is likely to be least precise? Why?

3. Describe a procedure for determining the total volume of your body.

## SAMPLE DATA SHEET:

Rectangular solid:

length:

width:

height:

mass:

calculated total surface area:

calculated volume:

calculated density:

calculated specific gravity:

Graduated cylinder:

diameter:

height:

calculated radius:

calculated volume:

calculated total suface
area of 100 ml of water:

Irregular solid:

measurements:

calculated volume:

## LABORATORY ACTIVITY 6:

## THE USE OF THERMOMETERS

### GENERAL INTRODUCTION:

The clinical thermometer is one of the physician's most important tools. It is used not only as an aid in the diagnosis of disease, but also to trace the course of a patient's recovery. Both fever (abnormally high temperature) and subnormal temperature have meaning to the physician. Whether the body temperature fluctuates periodically, increases, decreases, or changes in response to medication are all important things for the doctor to know. For this reason, the temperature of many hospital patients is taken periodically and a chart of the patient's temperature over time is kept at the foot of the bed.

A second type of thermometer, the laboratory thermometer, is also important in biomedicine, because many laboratory tests and chemical preparations depend to one extent or another on temperature. The most commonly used instrument for measuring temperature in the laboratory is the mercury thermometer. Although simple in design and easy to use, the thermometer is capable of providing valuable information quite quickly.

This activity is divided into two parts. The first part provides some guidance in the techniques of temperature measurement, while the second part will give some insight in the basic principles underlying thermometry and alternative ways of measuring temperature.

## PART I: USING THE MERCURY THERMOMETER

### INTRODUCTION:

Examine the thermometer. Notice that it is simply a glass tube that contains mercury. Most of the mercury is contained in a bulb at one end of the thermometer. This end is immersed in the material whose temperature is desired. The height of the mercury column is directly related to the temperature, which is read off the scale that is etched in the glass. Most laboratory thermometers are calibrated in degrees Celsius, although some include a Fahrenheit scale as well.

The thermometer is a useful tool only if one is sensible and careful while using it. It is easily broken and rather expensive to replace. If a thermometer should get broken, keep in mind that mercury is a toxic substance. If mercury gets on your hands, wash them afterward.

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MATERIALS:

- thermometer
- beaker, 250-ml
- beaker, 50-ml
- ring stand
- ring-stand ring
- wire gauze
- gas burner
- matches
- stirring rod
- ice
- table salt (5 teaspoonfuls)
- scoopula (or plastic spoon)
- clock with second hand

PROCEDURE:

1. Examine the temperature scale on the thermometer. What are the highest and lowest temperatures that can be read on the scale? (1)

2. What do you think would happen if the thermometer were placed in a liquid with a temperature of 300 °C? (2)

3. What would happen if the thermometer were placed in a liquid with a temperature of -35 °C? (3)

4. What is the temperature of the air in the room, to the nearest 0.5 °C? (4) Does the position of your eye have an effect on the value of the reading you obtain? (5)

5. Fill a 250-ml beaker with hot tap water. Immerse at least two inches of the bulb end of the thermometer in the water. Note the time to the nearest second as you do this. How long does it take for the mercury to equilibrate (to stop rising)? (6)

6. Hold the thermometer so that the bulb is located in the center of the water. Record the temperature to the nearest 0.5 °C. (7)

7. Hold the thermometer with the bulb touching the bottom of the beaker, and watch the mercury column closely. Does the height of the mercury column change? (8) How much of a temperature difference can be demonstrated in this way? (9)

8. Hold the thermometer with the bulb against the side of the beaker, and compare this temperature to the temperature in the center of the water. What do these results suggest about proper technique for measuring the temperature of a liquid? (10)

9. Suppose you wished to monitor the temperature of a liquid continuously over a long period of time. What part of the design of a laboratory thermometer

facilitates long-term temperature measurement? (11) What would you need to do to eliminate the need for holding the thermometer in the liquid for a long time? (12)

10. Immerse a finger in the warm water. Adjust the temperature of the water by adding hot or cold tap water until the warmest comfortable temperature is reached. (This is the temperature above which it is uncomfortable to leave the finger in the water.) Record this temperature to the nearest degree. (13)

11. Empty the beaker and re-fill it with ice. Add cold tap water until it reaches the top of the ice. Use the stirring rod to stir the icewater briefly. Determine the temperature of the icewater to the nearest 0.5 °C. (14) Why shouldn't the thermometer be used as a stirring rod? (15)

12. Add about a teaspoonful of salt to the icewater and stir for several seconds. Determine the new temperature. (16)

13. Empty and rinse out the beaker. Then add about 150 ml of hot tap water. Use a gas burner to heat the water until it is boiling. Determine the temperature of the boiling water. (17) Why doesn't the temperature of the water continue to rise as the boiling water is heated by the burner? (18)

14. Add four teaspoonfuls of salt to the boiling water and stir. When the liquid is again boiling, determine the temperature. (19)

## PART II: LIQUID AND GAS THERMOMETERS

### INTRODUCTION:

The mercury thermometer is based on the expansion and contraction of mercury in response to changes in temperature. But this property is not special to mercury; it is a characteristic of almost all materials. For certain special purposes, thermometers are used that contain other liquids or even gases, instead of mercury. This part of the activity is designed to give you some idea of how other thermometers work, or even how to make a thermometer, should you care to do so.

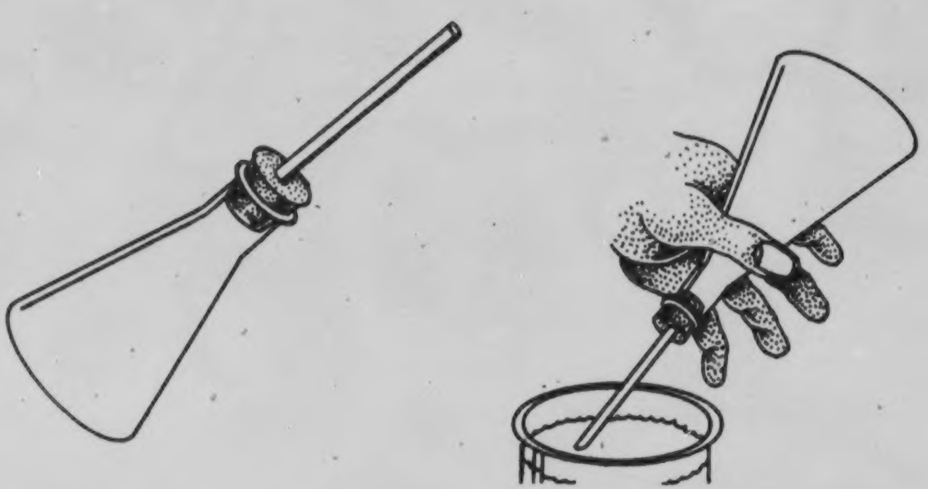
### MATERIALS:

- Erlenmeyer flask, 125-ml, with one-hole stopper
- length of glass tubing (at least 30 cm)
- beaker, 250-ml
- ring stand
- ring-stand ring
- wire gauze
- gas burner
- matches
- medicine dropper

PROCEDURE:

1. Insert the glass tubing into the stopper so that the end of the tube is flush with the small end of the stopper. Then stopper the flask tightly, as shown in the diagram.

2. Invert the flask and put the end of the glass tubing into a beaker partially filled with tap water.



3. Cup both hands around the flask, and watch the end of the glass tubing that is submerged in the water. Describe the results, and the apparent reasons for this behavior. (20)

4. If you had some way to cool the air inside the flask below room temperature, what would you expect to happen then? (21)

5. If you wanted to make a gas thermometer, it would be necessary to have the glass tube partly filled with water before it was submerged. Why? (22) Describe as fully as you can the other steps you would have to take in order to make a gas thermometer that would measure the temperature of the laboratory as it changed over time. (23)

6. Mercury freezes at about -39 °C and boils at about 357 °C. Describe some situations in which a gas thermometer would be useful, but a mercury thermometer wouldn't. (24)

7. Unstopper the flask and fill it to the brim with tap water. Do this in a sink or other large container, because you are going to spill a little water. Stopper the flask tightly as before. Use a dropper to fill the glass tube about half full of water.

8. Place the flask over a burner flame, and observe the water in the tube. Describe the result. (25) What important characteristics does this apparatus have in common with a mercury thermometer? (26)

9. How is a water thermometer limited in comparison to a mercury thermometer? (27)

10. Methyl alcohol thermometers are commonly used for the measurement of outdoor temperatures. Methyl alcohol freezes at about -94 °C and boils at about 65 °C. Discuss why methyl alcohol is an almost ideal fluid for outdoor temperature measurement. (28) Hint: it may be useful to convert the temperatures given to degrees Fahrenheit.

## LABORATORY ACTIVITY 7:

## USING THERMISTORS TO MEASURE TEMPERATURE

### GENERAL INTRODUCTION:

Suppose that you are working as a weather reporter, and that each day you need to know the temperature of 20 different towns in an area. Thermometers might be installed in each of these places. However, it would be quite a job to drive to 20 different towns every day in order to read the thermometers. This would take more time than it would be worth.

One way of solving this problem would be to use temperature-measuring devices called thermistors. The amount of electricity flowing through a thermistor depends upon temperature. The temperature is determined by measuring the electricity passing through the thermistor. An electric signal corresponding to the temperature can be transmitted miles away to a receiving station. In this way, the temperature of many locations may be monitored at the same time.

Thermistors have other advantages over thermometers. They are usually quite small and almost unbreakable. In addition, they respond more quickly to temperature changes. Also thermistors are better than thermometers for measuring the temperature of surfaces such as skin. On the other hand, thermometers have the advantage of being self-contained; they do not require an external electrical circuit in order to measure temperature.

Thermistors are used in many different fields. In medicine, they are being used more and more frequently to measure body temperature, as you will do in this activity. Other applications of thermistors include monitoring temperature in satellites, nuclear reactors, and in automobile engines.

As mentioned above, a thermistor requires a special electrical circuit. The device you will use for this purpose is called the Biomedical Instrumentation Package, more commonly known as the BIP. This instrument was designed especially for this course. It can be used to perform a number of laboratory tasks, and it will be used in many of the activities to come. Later in the course, you will learn how the parts inside the BIP work.

As with any electronic instrument, the BIP should be treated carefully. Even though its construction is fairly rugged, it should not be handled roughly or dropped. Spilled chemicals of any kind must be wiped up as promptly as possible. If treated with respect, the BIP can provide a wealth of valuable information.

It will allow you to do things that have never been done before in high school science courses.

## PART I: MEET THE BIP

### INTRODUCTION:

Since this is your first experience with the BIP, it will be necessary to spend a few minutes becoming familiar with its various parts. Don't worry about remembering all of the names just yet. They will come soon enough with experience.

### MATERIALS:

BIP

### PROCEDURE:

Refer to Figure 2 as you read the following description.

The different functions of the BIP are made possible by connecting various components, both inside and outside of the BIP. These connections are made by either plugging wires or electronic devices into the programming panel (11) on the rear of the instrument or the input jacks (1,8) on the front. The dials (2,4,6) provide data for a variety of applications. The slotted controls (3,5,7) are used to make slight adjustments when making certain measurements. Also on the front panel are a switch (9) and two diode lights (12). The BIP is turned on by plugging a line cord from a wall outlet into the line cord socket (10).

Each different BIP function requires different connections to the programming panel. An example of this programming is shown in Figure 1 below. The diagram indicates that wires connect pins H to K and T to U. In addition, a thermistor is attached between pins G and S. Although this is straightforward, it is easy to make programming errors when setting up the BIP. Much time can be saved by checking your wiring carefully before beginning any procedure utilizing the BIP.

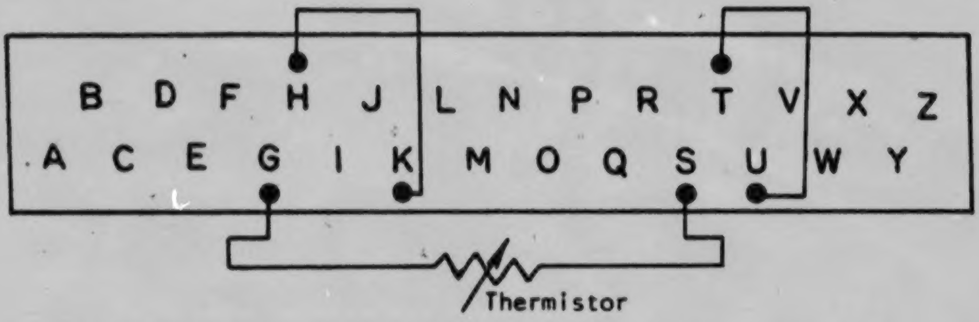
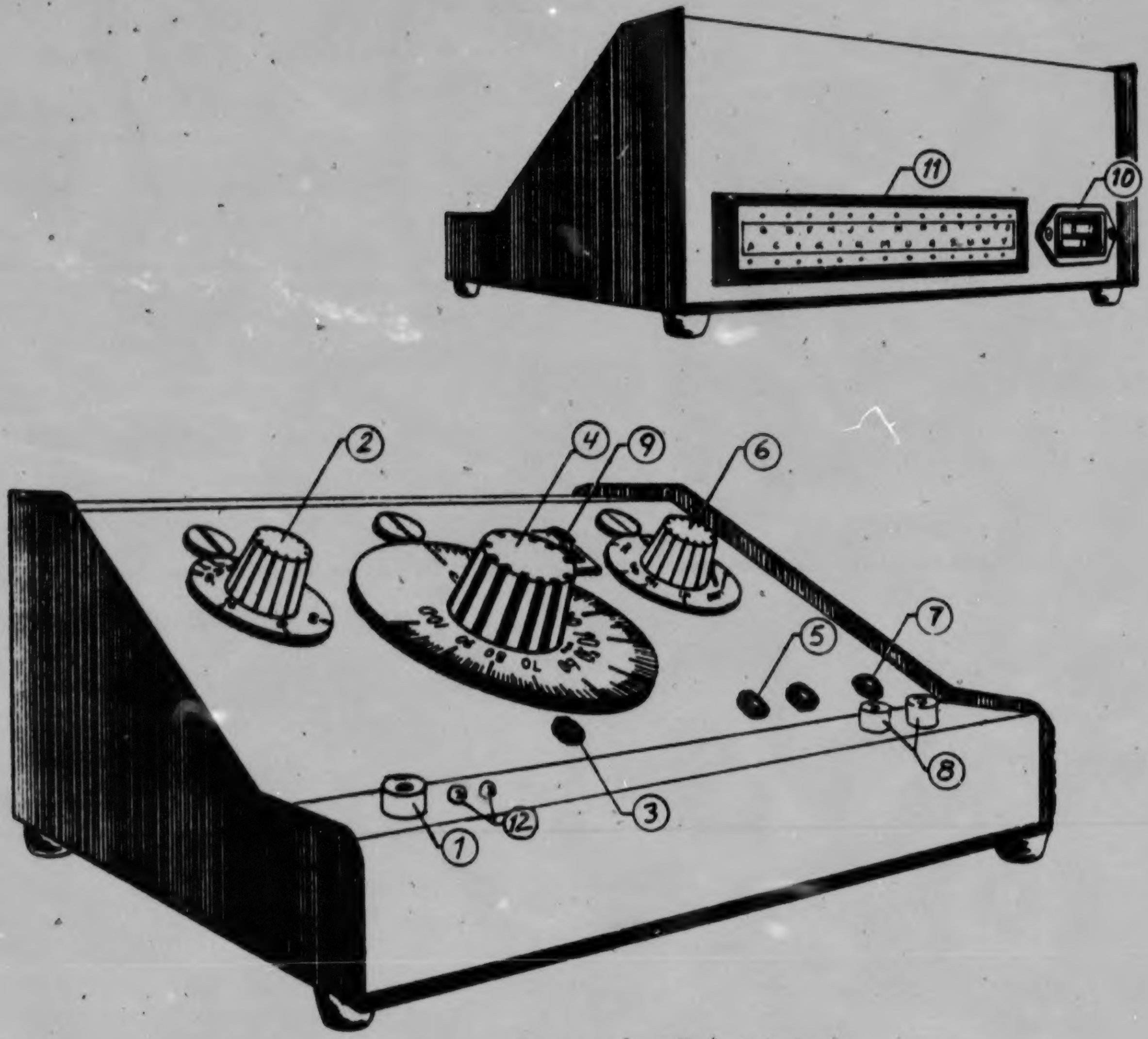


FIGURE 1: *Programming the BIP to Measure Temperature.*



1. pH input jack
2. amplifier gain control
3. variable voltage control
4. mA 0-100 dial
5. pH calibrate control
6. frequency control
7. pre-amplifier gain control
8. differential amplifier input jacks
9. frequency range switch
10. line cord socket
11. programming panel
12. diode lights

FIGURE 2: *The Parts of the BIP.*

## PART II: CONNECTING THE THERMISTOR TO THE BIP

### INTRODUCTION:

In this part of the activity, the BIP will be "programmed" to function as a device to measure temperature. This will entail connecting a thermistor component with wires ("programming wires") to the programming panel on the back of the BIP.

The following five cautions will be important to keep in mind every time you use the BIP.

1. Be sure that the line cord is disconnected until the program is known to be correct.

2. Use only 24-gauge solid wire with about 5 to 8 mm of insulation removed from either end. (See Figure 3.) If wire larger than 24-gauge is used, the pin may be sprung open too wide and will not grasp a wire used later.

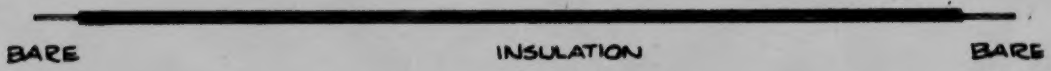


FIGURE 3: *A Programming Wire.*

3. Each wire end used in programming should be straight before it is pushed into a terminal.

4. Wire strippers should be used to remove about 6 to 8 mm of insulation from each end of the wire. The stripper should not cut into the wire itself; if it does, the wire will be weakened and may break off in the terminal. Proper adjustment of the "stop" on the stripper will prevent this.

5. Insert the wire by pushing the bare end straight into the pin connector (see Figure 4). If care is used, it should go in smoothly when you apply moderate pressure. If the wire end bends, cut off the bent part and re-strip the wire.

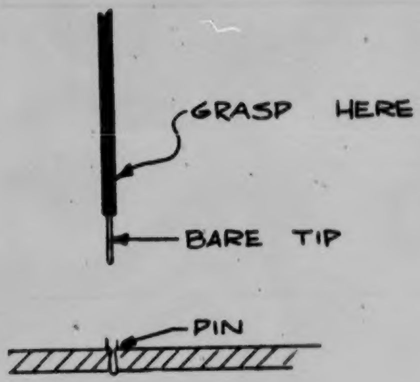


FIGURE 4:
*Inserting a Programming Wire Into a Pin*

MATERIALS:

- BIP
- thermistor component
- programming wire
- wire cutter-strippers

PROCEDURE:

1. Program the BIP as shown in Figure 1.

2. Check your programming to make sure it is correct. Only then plug in the BIP. The diode light that goes on indicates that the BIP is turned on.

3. Allow at least 10 minutes of warm-up time before beginning Part III--the longer the better. While you are waiting for the BIP to warm up, suspend the termistor in the air so that it does not touch any surface.

4. Turn the mA dial of the BIP and watch the diode lights. At some point during the rotation of the dial, one of the diode lights shuts off, while the other goes on. This point is called the null point. It indicates where a reading on the mA dial should be taken.

## PART III: CALIBRATING THE BIP

INTRODUCTION:

In Part II, you programmed the BIP so that it may be used to measure temperature. But the temperature readings cannot be read directly off the BIP. At the null point, the mA dial of the BIP is related to temperature as shown in the table on the next page. The relationship between mA readings and temperature has already been determined for each temperature value in the table. Your task, in Part III, will be to adjust the BIP so that the mA readings will correspond to the temperature readings given in the table.

MATERIALS:

- BIP with thermistor component (programmed)
- mercury thermometer
- screwdriver

PROCEDURE:

1. Read the room temperature in °C off of a mercury thermometer that is held near the thermistor. Once you have determined the room temperature, find the nearest temperature in the table that corresponds to this reading.

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## CONVERSION OF mA READINGS TO TEMPERATURE VALUES

| mA Reading | T (°C) | mA Reading | T (°C) | mA Reading | T (°C) |
|---|---|---|---|---|---|
| 100 | 38.1 | 79 | 32.1 | 58 | 24.1 |
| 99 | 37.8 | 78 | 31.7 | 57 | 23.7 |
| 98 | 37.6 | 77 | 31.4 | 56 | 23.2 |
| 97 | 37.3 | 76 | 31.1 | 55 | 22.8 |
| 96 | 37.1 | 75 | 30.7 | 54 | 22.3 |
| 95 | 36.8 | 74 | 30.4 | 53 | 21.8 |
| 94 | 36.5 | 73 | 30.0 | 52 | 21.3 |
| 93 | 36.2 | 72 | 29.7 | 51 | 20.8 |
| 92 | 36.0 | 71 | 29.3 | 50 | 20.3 |
| 91 | 35.7 | 70 | 29.0 | 49 | 19.8 |
| 90 | 35.4 | 69 | 28.6 | 48 | 19.3 |
| 89 | 35.1 | 68 | 28.2 | 47 | 18.7 |
| 88 | 34.8 | 67 | 27.8 | 46 | 18.2 |
| 87 | 34.5 | 66 | 27.4 | 45 | 17.6 |
| 86 | 34.2 | 65 | 27.1 | 44 | 17.0 |
| 85 | 33.9 | 64 | 26.7 | 43 | 16.5 |
| 84 | 33.6 | 63 | 26.2 | 42 | 15.8 |
| 83 | 33.3 | 62 | 25.8 | 41 | 15.2 |
| 82 | 33.0 | 61 | 25.4 | 40 | 14.6 |
| 81 | 32.7 | 60 | 25.0 | 39 | 13.9 |
| 80 | 32.4 | 59 | 24.6 | | |

2. Note the value in the column marked "mA reading" that is next to the room temperature value. Set the mA dial to this value. For example, if the room temperature is about 25 °C, the mA dial should be set at 60. If the room temperature is 22 °C, set the mA dial about half way between 53 and 54.

3. Once the mA dial has been correctly set for room temperature, insert a screwdriver into the slotted control of the voltage regulator (3). Turn the screwdriver until the diodes null. (Note: this adjustment is important. It is good technique to check the mA dial after setting the slotted control. The diode lights should null when the mA dial is set at the reading corresponding to the measured room temperature.)

4. The BIP is calibrated to read temperatures from 13.9 to 38.1 °C when the slotted control is correctly set. This control should not need further adjustment during the class period. Remove the screwdriver from the opening.

In order to measure a temperature using the thermistor, simply turn the mA dial until the diodes null. The temperature may be read from the table. For example, a mA reading of 81 indicates a temperature of 32.7 °C.

## PART IV: MEASURING SKIN TEMPERATURES

### INTRODUCTION:

In this part of the activity you will use the thermistor to measure skin temperature at various points on the body. The amount of heat given off by the skin is closely related to the amount of blood flowing in the vicinity and the rate at which it is flowing.

In clinical use, skin temperature measurements can provide valuable information on the location and severity of abnormalities in the circulation of the blood.

### MATERIALS:

BIP with thermistor component
cotton

### PROCEDURE:

1. To measure skin temperature, simply place the thermistor flat against the area that is being measured. Apply enough pressure to make a slight indentation in the skin. Place a small piece of cotton over the thermistor so that it is not exposed to the room air. Wait one minute for the temperature of the thermistor to equalize.

2. After one minute has elapsed, adjust the mA dial until the light diodes null. Record the mA reading and use the table to convert this reading to °C.

3. Measure the temperature of several locations on the body. Try the hands, arms, head and neck. After each measurement, record your results. A sample data sheet is provided.

4. Select a student who is willing to do a little exercising (see Step 5). Measure the temperature of the four locations indicated in Figure 5. Also measure the temperature of the subject's armpit. In this case, the thermistor can be held in place by having the subject close his or her arm on it.

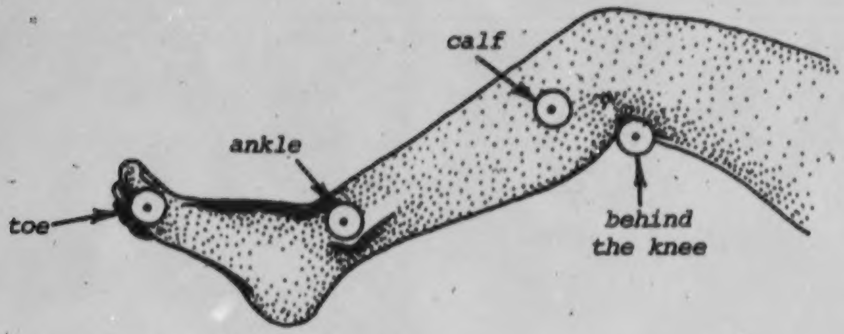

FIGURE 5: *Sites for Temperature Measurement Before and After Exercising.*

5. After completing Step 4, have the student exercise vigorously for several minutes. The person might run in place for five minutes or run up and down a flight of stairs five or six times.

6. After the exercise is complete, repeat Step 4.

DISCUSSION QUESTIONS:

1. Can you draw any general conclusions from comparing all the readings taken without exercise? Which locations gave the highest temperatures? Which the lowest temperatures? Is there a general trend?

2. How did the temperatures following exercise compare with those before exercise? Can you account for the changes?

3. If you found the temperature of a person's left hand to be significantly lower than the right-hand temperature, what might this indicate?

SAMPLE DATA TABLES:

| Location | mA Reading | Temperature (°C) |
|---|---|---|
| Palm of hand | | |
| Back of neck | | |
| Center of forehead | | |
| Cheek | | |
| Inner forearm | | |
| Shoulder | | |

|  | mA Reading |  | Temperature (°C) |  |
|---|---|---|---|---|
| Location | Before Exercising | After Exercising | Before Exercising | After Exercising |
| Behind knee |  |  |  |  |
| Calf |  |  |  |  |
| Ankle |  |  |  |  |
| Toe |  |  |  |  |
| Armpit |  |  |  |  |

# LABORATORY ACTIVITY 8:

# OXYGEN

## INTRODUCTION:

As was discussed in Section 8, of the gases that make up our air only one--oxygen--is essential to our health and life. Approximately 21 per cent of the air is oxygen.

In this activity you will have an opportunity to produce pure oxygen gas and to test its behavior in a variety of ways. You will compare its properties to the properties of air and then, in Laboratory Activity 9, to the properties of another gas that is important to respiration--carbon dioxide.

The activity consists of three parts. The first part involves the generation and collection of oxygen in several test tubes. In the second part, the oxygen samples will be subjected to a number of tests. The third part of the activity involves observations of the interaction of oxygen and steel wool over a period of several days.

## PART I: GENERATING OXYGEN GAS

### MATERIALS:

- potassium chlorate
- manganese dioxide
- large test tube with 1-hole stopper
- glass tubing, at least 5 cm long
- rubber tubing, at least 75 cm long
- ring stand
- ring-stand clamp
- 6 test tubes, 16 x 125 mm, with stoppers
- gas burner
- matches
- balance
- scoopula
- spatula
- pan of water (or large beaker)
- steel wool
- glass-marking pencil

### PROCEDURE:

1. Place approximately 5 g of potassium chlorate in a large test tube. Using a spatula as a measure, add one scoop of manganese dioxide to the potassium chlorate. Only a tiny amount of this material is needed; the results are not improved by adding more than the specified amount.

2. Clamp the test tube to a ring stand and assemble the apparatus as shown in Figure 1 on the next page. Be certain that the stopper is pushed tightly into the mouth of the test tube. The rubber tubing should be long enough

to reach the pan of water. This tubing will carry oxygen gas which is collected in test tubes immersed in the water. When these test tubes are filled with water and inverted in the pan, oxygen gas may be bubbled in from below.

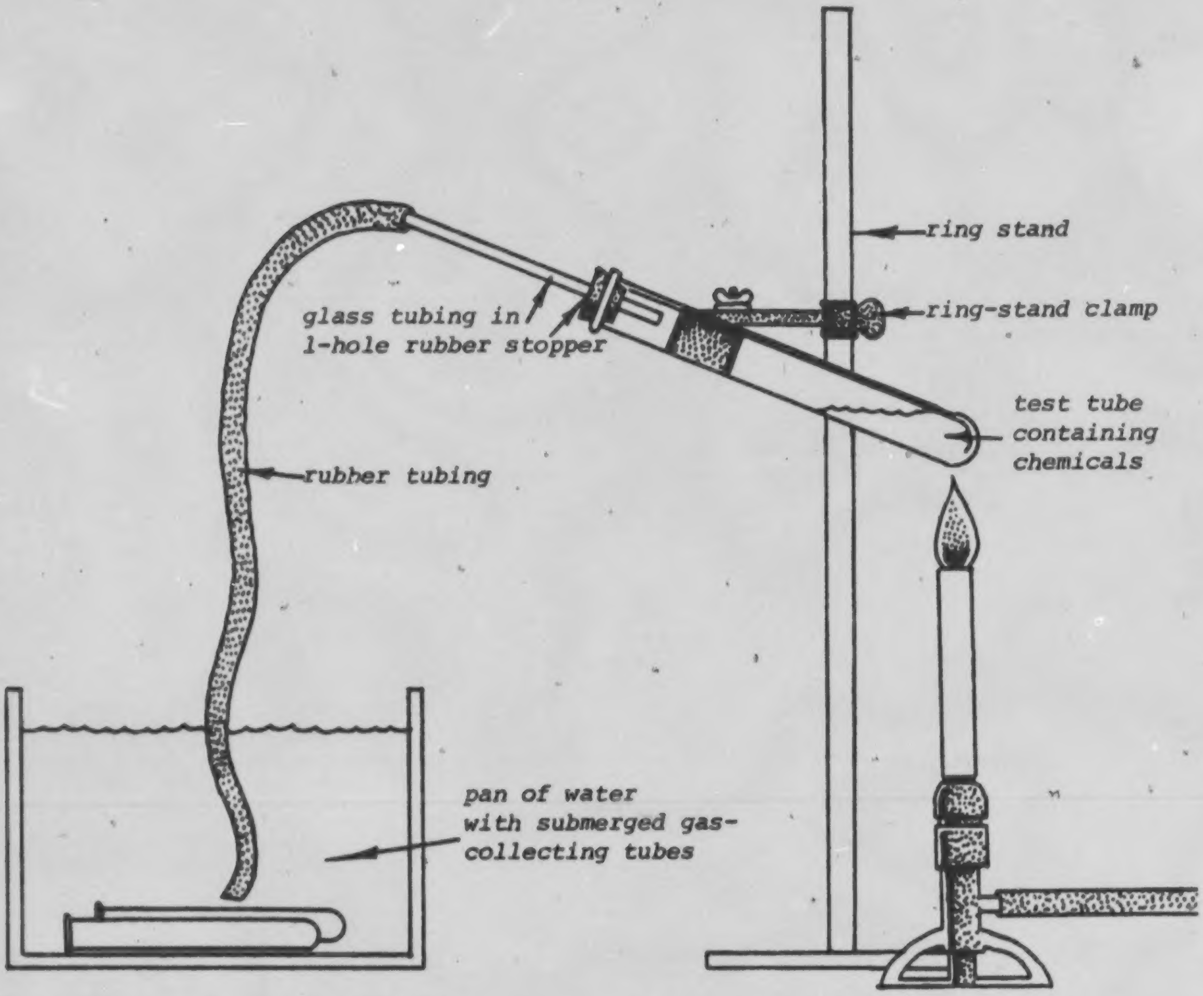

FIGURE 1: *Gas-generating apparatus.*

3. Take a small wad of steel wool and stuff it loosely into the end of a small test tube. The steel wool should not be tightly packed, but it is essential that it stay in the end when the tube is inverted and jostled slightly.

4. Place the test tube prepared in Step 3 and five other small test tubes in the pan of water. These are the gas-collecting tubes. They should be submerged so that all of the air in the tubes bubbles out.

5. Light the burner. After adjusting the flame, place the burner under the test tube containing the mixture of chemicals.

6. As the test tube heats up, gas begins to bubble into the water. The first few bubbles are mostly air, and do not need to be collected. Now watch the chemicals inside the test tube. As the temperature rises, the potassium chlorate melts. When the chemicals start bubbling vigorously, begin collecting gas in the submerged test tubes as shown in Figure 2.

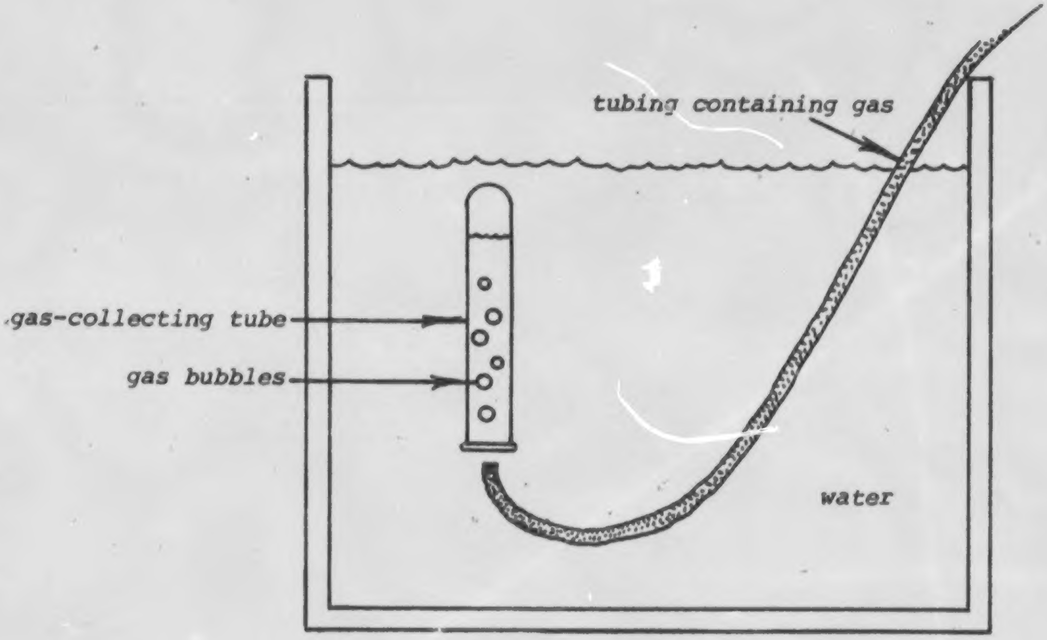

FIGURE 2: *Filling a tube with gas.*

7. After each tube is filled with oxygen, quickly place a stopper over its mouth and transfer the rubber tubing to another test tube. It is very important to keep the gas collecting tubes submerged until they are stoppered. Air should not be allowed to mix with the oxygen inside the test tubes. Don't worry about a small amount of water in each tube.

8. Continue collecting oxygen until all test tubes have been filled. Reduce the heat if the bubbling becomes too vigorous. When enough oxygen has been collected, turn the burner off.

9. Label the oxygen tube containing steel wool with an "O" for oxygen, and put it aside for use in Part III.

## PART II: PROPERTIES OF OXYGEN GAS

### MATERIALS:

5 oxygen samples (from Part I)
phenol red
medicine dropper
wooden splints
matches

### PROCEDURE:

1. Light the end of a wooden splint and then blow it out. While the splint is still glowing, uncap one of the gas tubes and quickly thrust the splint into the gas. Record your observations.

2. Uncap one of the gas tubes and quickly add a dropperful of phenol red. Recap the tube and shake for a few seconds. Record your observations. Rinse out the tube.

3. Repeat Step 2 with a tubeful of air. Record your observations.

4. Break a wooden splint so that it can fit completely inside one of the gas tubes. Light the splint, blow it out and thrust the glowing splint into a tube containing oxygen gas. Re-stopper the tube. When the splint has burned out, repeat Step 2 on this tube. Record your observations.

## PART III: THE INTERACTION OF OXYGEN AND AIR WITH STEEL WOOL

### MATERIALS:

oxygen sample with steel wool
  (from Part I)
2 test tubes, 16 x 125 mm
beaker, 150-ml
graduated cylinder, 10-ml
millimeter ruler
medicine dropper

### PROCEDURE:

1. Place about 100 ml of tap water in a 150-ml beaker.

2. Invert the oxygen sample containing the wad of steel wool and place it in the beaker. With the mouth of the tube submerged, remove the stopper. Leave the tube resting against the side of the beaker.

3. Prepare a second test tube with a wad of steel wool just as you did in Step 3 of Part I. Label this tube "A" for air.

4. Fill this tube with water to the brim. Using a 10-ml graudate, measure the amount of water in the test tube to the nearest 0.5 ml. This volume is

equivalent to the volume of air that the tube contains; that is, the total volume minus the volume of the steel wool. (Note: if the tube holds more than 10 ml, the graduate must be filled to the 10-ml mark and then emptied. The rest of the water in the test tube is measured with the graduate, and this volume is added to 10.0 ml to give the total volume.) Record the result in your data sheet.

5. Invert the test tube containing the steel wool and air and place it in the beaker of water. Allow it to rest against the side of the beaker. If you are careful, air will not bubble out.

6. Take a third test tube (empty) and place it inverted in the beaker. Place the beaker in a location where it will be safe for several days.

7. Observe the test tubes daily. Record any significant changes as they occur, including a comparison of the water levels in the three tubes. When the water level stops rising in the tube containing air and steel wool, make a final comparison of the three tubes. The remainder of this procedure applies only to the tube with air and steel wool. The other two tubes may be removed from the beaker.

8. Measure the height of the water level in the tube by placing one end of a millimeter ruler in the water. (If the "zero" is not at the end, it doesn't matter.) Hold both the ruler and the test tube as vertically as possible resting against the bottom of the beaker. Make your reading through the side of the beaker. Record the measured height.

9. Remove the test tube and ruler from the beaker. Hold both vertically against the desk top. Use a glass-marking pencil to mark the position on the side of the test tube of the height you recorded in Step 8.

10. Determine the volume of gas that remained in the tube by filling the tube with water to the level marked in Step 9. (Use a dropper to adjust the water level.) Measure the volume of water with a graduated cylinder. Record this volume on your data sheet.

11. Find the volume of gas consumed by subtracting the volume obtained in Step 10 from the initial volume obtained in Step 4. Record the difference.

12. Express the volume of gas consumed as a percentage of the initial volume.

$$\frac{\text{ml gas consumed}}{\text{ml gas initially}} \times 100 = \text{percentage of gas consumed}$$

## DISCUSSION QUESTIONS:

1. Compare oxygen and air in terms of their ability to support burning or combustion.

2. Compare oxygen, air and the gases resulting from the burning of a wooden splint in terms of their effect on phenol red.

3. Is oxygen the gas that is produced when a splint is burned? Explain your reasoning.

4. In Part III, explain why a tube containing nothing but air was placed in the beaker.

5. What happened to the gas consumed in Part III? What gas was consumed?

6. What is the significance of the percentage obtained in the final step of Part III?

## SAMPLE DATA SHEET:

Part II Observations:

oxygen + glowing splint:

oxygen + phenol red:

air + phenol red:

phenol red + gas resulting from burning splint in oxygen:

Part III Data:

initial volume of gas:

final height of water level:

final volume of gas:

calculated volume of gas consumed:

calculated percentage of gas consumed:

Part III Observations:

## LABORATORY ACTIVITY 9:

## CARBON DIOXIDE

### INTRODUCTION:

In Section 9-2, we discussed the importance of carbon dioxide gas in human respiration. Unlike oxygen, carbon dioxide is a gas that is removed from the body during respiration. Although the air we breathe in contains only about 0.03 per cent carbon dioxide, the air we breathe out is about 5 per cent carbon dioxide.

In Laboratory Activity 8, you generated oxygen gas and tested its properties in several ways. In this activity you will generate carbon dioxide gas and perform some of the same tests on it.

### MATERIALS:

marble chips
hydrochloric acid
limewater
ammonia water
phenol red
gas-generating apparatus (from Laboratory Activity 8)
4 beakers, 50-ml
6 test tubes, 16 x 125 mm, with 3 stoppers
balance
medicine dropper
scoopula
glass-marking pencil
wooden splint
matches

### PROCEDURE:

1. Carbon dioxide may be produced with the apparatus that was used for collecting oxygen gas. A burner is not needed for this procedure. Assemble the apparatus as you did in Laboratory Activity 8, making sure that the equipment is clean. Submerge six gas-collecting tubes in the pan of water.

2. Place approximately 25 ml of tap water in a labeled 50-ml beaker. Set the beaker near the pan of water holding the gas-collecting tubes, but not where it will get in your way while collecting gas samples.

3. Repeat Step 2 using first limewater and then ammonia water, so that you have three labeled beakers of liquid near the gas-collecting pan.

4. Place approximately 5 g of marble chips in the large test tube where the gas will be generated.

5. Obtain approximately 25 ml of hydrochloric acid in a clean beaker. Note: Hydrochloric acid is a dangerous substance only if it is handled improperly.

Avoid splashing or spilling this chemical, and keep it away from the eyes. If hydrochloric acid comes in contact with the skin, do not panic. Walk to the nearest sink and rinse the skin thoroughly with water.

6. Add the acid to the marble chips. Quickly place the stopper connected to the rubber tubing over the mouth of the test tube. Gas should immediately begin bubbling out of the rubber tubing in the pan of water. After 30 seconds have passed, most of the gas bubbling out of the tubing should be carbon dioxide.

7. Collect and stopper the first three gas samples just as you did in Laboratory Activity 8. Be sure the stoppering is done underwater, so that no air mixes with the carbon dioxide.

8. When the fourth tube is filled with gas, place a finger over the mouth of the tube, instead of using a stopper to trap the gas. With the tube inverted, remove it from the pan and place it in the 50-ml beaker containing tap water. Remove your finger from the mouth of the tube only after it is submerged, so that no gas escapes. Record the time.

9. Record your observations after 1 minute, 5 minutes and 15 minutes. (Go on to Step 10 while waiting to make your observations.)

10. Repeat Steps 8 and 9 using the last two test tubes and the beakers containing limewater and ammonia water. Rinse your hands briefly after touching the limewater and ammonia water. Perform the remaining steps as time permits in between or after the observations made on the beakers.

11. Light a wooden splint and blow it out. Uncap a gas sample and quickly thrust the glowing splint into the gas. Record your observations.

12. Uncap a gas sample and quickly add a dropperful of phenol red. Recap the tube and shake for a few seconds. Record your observations.

13. Rinse out the tube used in Step 12. Add a dropperful of phenol red. Exhale as completely as possible into the tube (allowing air to escape as you blow in), and then quickly cap it. Shake the tube and record your observations.

DISCUSSION QUESTIONS:

1. Compare carbon dioxide and oxygen in terms of the glowing splint test.
2. Discuss the ability of carbon dioxide to support combustion.
3. Compare carbon dioxide, oxygen and air in terms of the phenol red test.
4. Explain the result of performing the phenol red test on exhaled air.

5. Attempt to explain your observations on the tubes of carbon dioxide in the beaker of tap water and the beaker of ammonia water.

6. What do you think might be the result of performing the limewater test on pure oxygen? (Remember that air is about 21 per cent oxygen.)

7. Use all of your observations in this activity and your observations in Part II of Laboratory Activity 8 to write a two-part general statement on the properties of oxygen and the properties of carbon dioxide.

SAMPLE DATA SHEET:

| Observations after: | 1 min | 5 min | 15 min |
|---|---|---|---|
| carbon dioxide over tap water | | | |
| carbon dioxide over limewater | | | |
| carbon dioxide over ammonia water | | | |

Other observations:

carbon dioxide + glowing splint:

carbon dioxide + phenol red:

exhaled air + phenol red:

# LABORATORY ACTIVITY 10:

# EXAMINATION OF THE RESPIRATORY SYSTEM

## GENERAL INTRODUCTION:

In the Student Text, we have begun to examine the respiratory system--how it functions and some of the conditions that can affect the process of respiration. At the outset of this study, it would help to have a clear mental image of the human respiratory system and its parts. We can and will provide you with pictures, but pictures are a poor substitute for the real thing.

In this activity you will have an opportunity to study a set of beef or sheep lungs and trachea. The respiratory systems of such animals are much like our own. In fact, a great deal of medical and biological research is done on animals because of such similarities.

The activity will include a study of the general structure of the respiratory system and its relation to other systems, dissection of the bronchial tubes, determination of the density of different tissues and microscopic investigations.

A thorough analysis of a laboratory specimen involves concentration and full use of the senses. For example, people who work in a hospital laboratory are often asked to report on a sample of tissue taken from a diseased patient. Two different investigators might study the same tissue and perceive somewhat different things. The survival of the patient may well depend on how well the investigator uses his or her senses to come to a real understanding of what is being observed.

In this activity, and in all the activities to follow, try to use your senses fully. First and foremost is the sense of vision. Take a penetrating look at everything. Second is the sense of touch. Feel the different tissues to see how smooth, rough, rigid or elastic they are. Third, the sense of small can be useful. At least, it should tell you whether the lung sample is fresh or spoiled. Fourth is the sense of hearing. Most biological specimens are not exactly noisy, but you can learn something about the lungs by squeezing lung tissue and listening. Fifth is the sense of taste. You should not taste anything in the laboratory unless instructed to do so. (There will be very few activities that will involve tasting things.) Last is "common sense," which is not really a sense at all. Use your common sense and good judgment in the laboratory. Sometimes it may help to use a slightly different technique than the one described to make an observation.

This activity is too long to be completed in one period. When you clean up between periods, be sure to save specimens that will be needed later. Dispose of any unwanted tissues either by placing them in a plastic bag or as directed by your instructor. Do not throw any solids into the sink. Be sure to clean all dissecting tools, table tops and other equipment well, because fresh materials spoil easily. Otherwise, your sense of smell may rebel during later lab periods.

## PART I: APPEARANCE OF THE LUNG

### INTRODUCTION:

Before doing any dissection, the external appearance of the lung should be examined. Once dissection is started, the lungs will never be the same. Organs, such as lungs, function in association with other organs. In Part I, you will study the lungs and the surrounding tissues and organs.

In dissecting a specimen, it is helpful to have some way of orienting the specimen in terms of the whole organism. What part of the specimen is nearest to the head? What part is nearest to the front? Biologists have some useful terms to describe the parts of the organism as indicated in Figure 1. The region near the head is called the "anterior." The opposite end is called the "posterior." The back of an organism is referred to as the "dorsal" side and the front is the "ventral" side. You will need these terms to describe the lungs.

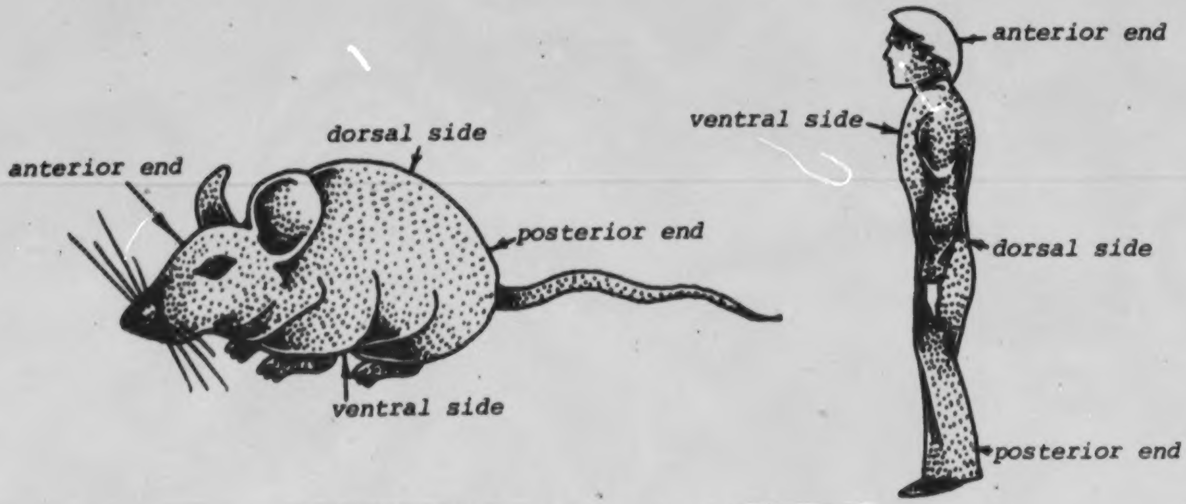

FIGURE 1: *Terms used to describe the orientation of a specimen.*

### MATERIALS:

beef or sheep lungs, fresh, in large tray or pan
tube, 4 cm diameter, at least 20 cm long (rubber, plastic or glass)

PROCEDURE:

1. Examine the entire specimen. Locate the trachea. The trachea is a long hollow rigid-walled tube with a diameter of 4 to 5 cm. Note that the trachea is connected to two large masses of tissue. These are the lungs. The trachea is nearest to the anterior end of the lungs.

2. The lungs are divided into subsections called lobes. How many lobes are there in each lung?

3. To distinguish the ventral side of the lungs from the dorsal side, do one of two things. Either have someone strong lift the specimen by the trachea or else roll the specimen over. In either case, look at both sides of the trachea. Note the rings that go around it. On the dorsal side, the rings are not continuous, but are separated by a groove which runs the length of the trachea (see Figure 2). On the ventral side, there is no groove and you should be able to find the cavity where the heart was. This cavity is lined with fatty tissue.

4. Observe the dorsal side of the trachea. In the groove, there may be a dark red muscular tube about 2 cm in diameter. This tube is the esophagus, a food tube, which connects the mouth with the stomach. Examine the muscular walls of the esophagus. How many layers of muscle can you see? How are the muscles arranged? (Are they around the tube or are they up and down the length of the tube?) Look inside the esophagus for remnants of food. What kind of food might you expect to find in a cow or sheep?

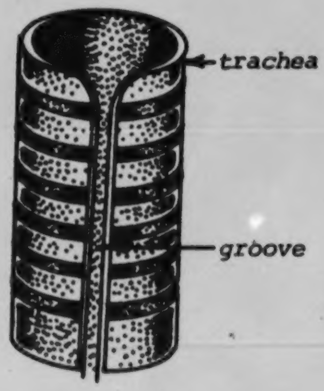

FIGURE 2:
*Trachea--dorsal view.*

5. Turn the specimen over so that the ventral side is up. Look in the cavity where the heart was. Examine the surface of this cavity below the fat for a thin, almost transparent membrane. This is the pericardium, a bag which surrounds the heart. Search in the masses of fat and the surface of the pericardium for the ends of tubes which vary from .5 to 3 cm in diameter. These tubes are blood vessels which carry blood to and from the heart, the lungs and the body. In preparing the specimen, these vessels were cut away from the heart and/or the lungs. Examine the walls of the tubes for thickness. Are these tubes elastic? Pull on them and see whether they are. Try to poke your finger through one of the larger tubes to see how tough it is.

6. Examine the lungs. Feel the surface of the lung and then pinch it. What kind of texture did you find? Was it what you expected? Note the color of the lungs in their deflated state.

7. To inflate the lungs, insert some tubing into the trachea. Hold your hand around the trachea to seal the space between the trachea and the tube. Put your mouth to the end of the tube and blow. You may need to blow several times before parts of the lung begin to inflate. If the lung does not inflate, look for an air leak in the lungs. Try to seal off the leak by pinching the lung together on both sides of the leak. Also, it may be possible to compensate for the leak by inserting the tubing farther into the trachea.

8. Unless there is an air leak, an inflated lung will tend to remain inflated for a time. If necessary, you can blow the lung up a second time. What color is the inflated lung? Touch the inflated lung and then pinch it. Compare the color and texture of the inflated and deflated lung.

9. If there is no time to continue, be sure to store the lungs in a refrigerator.

## PART II: DISSECTION--TISSUES OF THE RESPIRATORY SYSTEM

### INTRODUCTION:

Studying the whole lung or the whole trachea as you did in Part I is one way to learn about how the system works. Another way is to study the tissues and cells that compose these organs. Both approaches are valuable and provide different information. In Part II, you will dissect out some different kinds of tissue that contribute to the functions of the respiratory system. The tissues will be saved for other investigations in Parts III and IV.

### MATERIALS:

beef or sheep lungs
scalpel or razor blade (single-edged)
Petri dish

### PROCEDURE:

1. The air that we inhale passes through the trachea and later enters the lungs. How are the two organs connected? In order to answer this question, it will be necessary first to remove all excess fat and other tissue from the cavity where the heart was. Use a scalpel or single-edged razor blade with care to cut

away the pericardium and all the fat attached to it. Be careful not to cut any lung tissue. Save the fat for use in Parts III and IV. It may be stored in a refrigerator in a labeled Petri dish.

2. Follow the trachea down toward the lungs with your fingers. A few centimeters posterior to where the lungs begin, feel for the branching of the trachea. These branches are the bronchi. Use a scalpel to cut through the lungs to expose the region where the branching begins. Make short cuts and dissect slowly. This will allow the tissues to fall to the sides exposing the region of your next cut. Dissect the lung tissue away from one of the bronchi. As you do this, continue feeling for the bronchus and cutting away lung tissue, barely exposing the bronchus without cutting through it. As you work your way down the bronchus, notice the colorless tubes which lie next to it. These tubes are blood vessels with little or no blood in them. Notice how these vessels continue to branch into smaller and smaller tubes.

3. Continue your dissection until you get to what seems like the end of the bronchus. Actually this tube does not end abruptly. It subdivides into smaller and smaller tubes. What is the advantage of all this branching for the respiratory system? As you dissect, you may notice that you have cut through some small tubes which maintain their round shape. These smaller branches of the bronchus are called bronchioles. Why do these tubes maintain their round shape, Feel them for a clue. You may notice some other tubes near the bronchioles which are collapsed. What kind of tubes are these?

4. Look at the surface of the lung along the edge of a cut. Note the thin membrane which surrounds the outer surface of the lung. This membrane is the pleura. The narrow space beneath it is the pleural space.

5. Feel and observe the trachea. Note that it has hard rings alternating with softer tissue. The substance that makes the rings hard is known as cartilage.

6. Dissect out a ring of cartilage from the trachea. To do this, feel for the soft space between the first two rings at the anterior end of the trachea. Insert the scalpel blade in the soft tissue and cut through on both sides of the ring. Cartilage is soft enough to be damaged by the blade, so work carefully. Note the layer of tissue which lines the inside of the trachea. This is the "epithelium." Use the scalpel carefully to cut away some of the epithelium adjacent to the cartilage. This will take some patience and skill, but learning these techniques is worth the effort.

Save the cartilage for Parts III and IV. The cartilage may be wrapped in a paper towel and stored either at room temperature or in the refrigerator. Save the epithelium for Part IV. This tissue may be placed in the same Petri dish as the fat and frozen. If Part III will be done immediately, it will only be necessary to store the epithelial tissue. What are the other two kinds of tissue you have collected?

## PART III: DENSITY OF TISSUES

### INTRODUCTION:

In Part II, several different tissues were collected. Tissues differ in several ways. In this activity, you will determine the density of these and other tissues to see whether tissues differ in density. The techniques for determining density have already been covered. If necessary, review Laboratory Activity 4 before proceeding.

### MATERIALS:

beef or sheep lung

beef or sheep lung tissues (from Part II)

balance

graduated cylinder, 100-ml

scalpel or razor blade

beaker, 50-ml

### PROCEDURE:

1. In this part of the activity, you will compare the density of the following tissues.

   a. cartilage

   b. fat

   c. lung (actually a mixture of tissues)

   d. esophagus, if available (this is mostly muscle tissue)

2. It will be assumed that you know how to determine density by the immersion technique. For each determination, you will need a chunk of tissue(s) approximately 2 cm in size. Use a weighed beaker as a tare. (If paper were used as a tare, the pan of the balance might get wet.) The same beaker may be used to determine each of the masses if it is cleaned and dried between weighings. Use a graduated cylinder to determine volumes.

3. Remember that it is necessary to immerse an object completely in order to determine its volume. If any of the tissues float, figure out a way to immerse them without significantly changing the volume reading.

4. Record all data on a data sheet as you have done before. Determine and record the various tissue densities. Within the limits of the precision of your measurements, do all tissues appear to have the same densities?

5. Objects denser than water will sink, while objects less dense than water will float. Did the lung sample sink or float?

6. Squeeze the section of lung tissue between two fingers and listen. Do you hear something. It is not necessary to squeeze all the air out. Put the section back in the water. Does it still float? Is it more dense than before you squeezed it? Why?

7. Save the tissues for use in Part IV. Place them in a Petri dish and freeze.

## PART IV: MICROSCOPIC EXAMINATION OF TISSUES

### INTRODUCTION:

Thus far you have considered parts of the respiratory system on two levels--the organ and tissue levels. But recall that tissues are composed of large numbers of cells. In Part IV, you will use a microscope to examine the cellular substructures of the tissues you have collected. This will require some prior skill with the microscope, so Part IV has been divided into four parts as follows.

A. Parts of the Microscope

B. Making Microscope Slides

C. Use of the Microscope

D. Microscopic Examination of Tissues

### MATERIALS:

- compound microscope
- 9 to 11 glass slides
- 9 to 11 coverslips
- medicine dropper
- lens paper
- Kimwipes or other tissue
- scissors
- two dissecting needles
- small piece of newsprint from the classified section
- small piece of colored picture from a magazine
- small piece of cotton
- samples of tissue (from Parts II and III)
- methylene blue stain
- forceps
- Petri dish half (either top or bottom)

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## A: PARTS OF THE MICROSCOPE

INTRODUCTION:

The microscope is a valuable tool for a scientist. It enables him or her to see tiny structures which are not visible to the unaided eye. Microscopes are expensive tools which should always be used with proper technique to avoid damaging any of the parts. With proper care and use, the microscope will provide a great deal of insight into the fundamental nature of life.

PROCEDURE:

1. In order to use the microscope to study the tissues collected in Parts II and III, you will need to know the names of the important parts of the instrument. Obtain a microscope and, as you read the procedure, study the diagrams and find the corresponding parts on the microscope you have.

2. Find the parts of your microscope which are labeled in Figure 3. (Your microscope may not be identical to the one shown in Figures 3 to 6.)

a. A is either a mirror that reflects light (as in the drawing) or a small electrical lamp. The mirror or lamp is adjusted to send light up through the material being observed.

b. B is the diaphragm which controls the amount of light that passes through the material being observed.

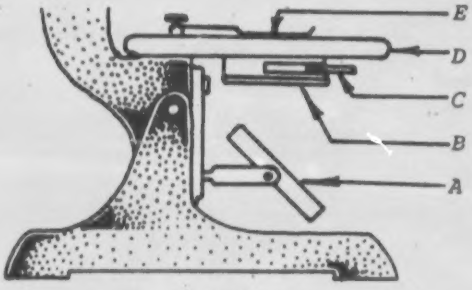

FIGURE 3

c. C is either a lever or dial which can be moved to adjust the amount of light passing through the diaphragm. The amount of light is a very important factor in the optimal use of a microscope. Try moving the lever gently or changing the dial setting. Later we will consider how such adjustments can increase or decrease the amount of light.

d. D is the stage. Glass slides containing specimens for viewing are placed on the stage. Note the hole in the center of the stage. What is it for?

e. E is a clip. There may be one or two clips on your microscope. The clip serves to hold slides in place. It is a good idea to secure a slide with at least one clip especially when you share a microscope. This will prevent the slide from being moved accidentally. If a slide is moved out of position, it may be very difficult to relocate what you were looking at.

3. Note Figure 4. Find the labeled parts on your microscope.

a. F is called the low-power objective. It is marked 10X. At the end of this metal tube is a lens which will magnify an object by a factor of 10.

b. G is called the high-power objective. It is marked 40X or 43X. At the end of the metal tube is a glass lens which magnifies things about forty times their actual size.

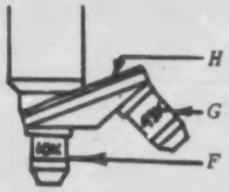

FIGURE 4

c. You may have a third objective or lens marked 4X. This lens magnifies things four times and is sometimes called a scanner. It is used to study larger objects or to position a slide on the stage so that the object to be studied is centered.

d. H is called the revolving nosepiece. The objectives are attached to the nosepiece. To change the magnification, the nosepiece is rotated placing a different objective over the hole in the center of the stage. Revolve the nosepiece, but be careful that the high-power objective does not hit the stage. The glass lens in an objective is very expensive.

When the objective is in its proper position, you should feel a "click" as it snaps into position. Always be sure the objective is "clicked" in its proper position before trying to use the microscope. Otherwise, no light will pass through the lens.

5. Study Figure 5. Find the labeled parts on your microscope.

a. J is the fine adjustment. It is a small knob. As the fine adjustment is turned, the objective is moved closer to or farther from the slide. This allows you to focus.

b. K is the coarse adjustment. It is a larger knob. Turn the coarse adjustment slowly, taking care that the objective does not hit the stage. Note how the objective moves with respect to the stage. Record on your data sheet whether clockwise or counterclockwise turning (as you face the knob) moves the objective closer to the stage. Now move the knob marked J in Figure 5. It should be clear why J is called the fine adjustment and K is called the coarse adjustment.

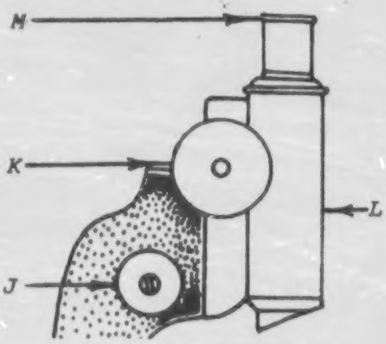

FIGURE 5

c. L is the body tube. The body tube is located directly above the objectives. To look through the body tube, find the lens at M.

d. M is the eyepiece. The eyepiece has a glass lens which also magnifies. Look for a marking on the eyepiece which indicates its magnification power. Look through the eyepiece. If the mirror or lamp is adjusted properly and an objective is "clicked" into position, you whould see a circle of light. This circle is called the field. Your eyepiece may have a pointer which looks like a needle sticking half way across the field of view. Some pointers can be moved by rotating the eyepiece. A pointer helps two people to discuss the position of what they are looking at through the microscope.

5. Figure 6 shows the entire mciroscope. It is a composite of Figures 3 to 5 and also shows two additional parts.

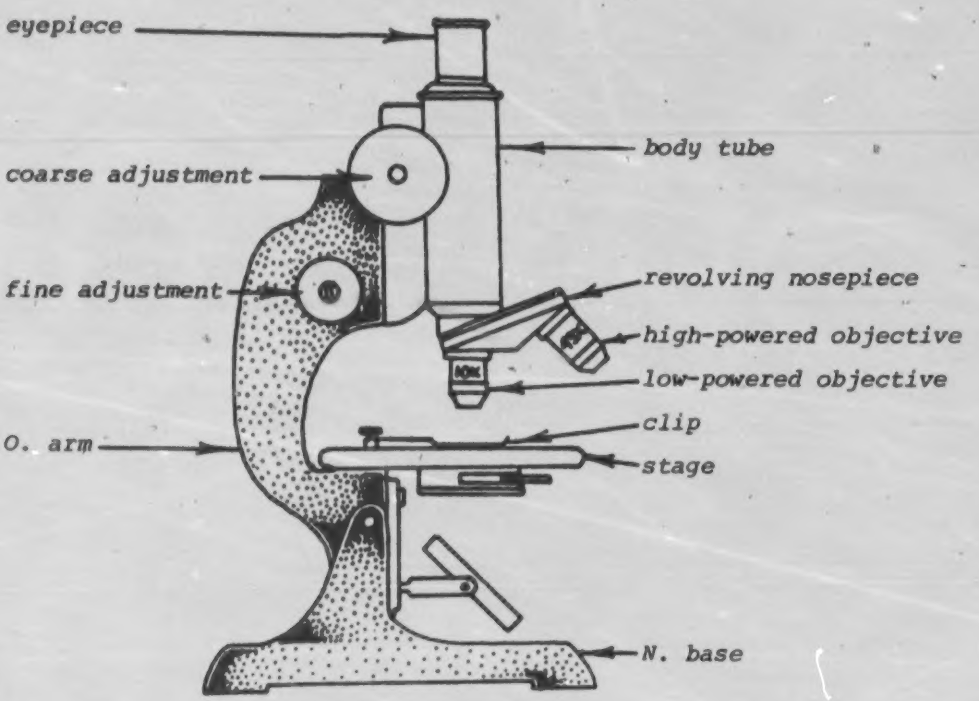

FIGURE 6

a. N is the base of the microscope. It should be placed on the table about 4 to 6 cm from the edge of the table with part O facing you.

b. O is the arm. When you carry the microscope, grasp the arm with one hand and place the other hand under the base. Always use TWO HANDS to carry a microscope! Dropping the microscope would be a minor calamity.

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c. Connecting the base and the arm there may be a screw which allows the microscope to be tilted. It is best to leave the microscope in a full upright position (no tilting) because many slides will have solutions on them which will run off if the stage is tilted, and because tilting unbalances the instrument and makes it unstable.

## B: MAKING MICROSCOPE SLIDES

### INTRODUCTION:

Besides using the microscope to study "prepared" slides that are available commercially, you will have occasion to make your own slides. The most common preparation you will make is called a "wet mount" because it involves liquids. In this part of the activity, this technique will be explained. Later you will have a chance to view the slides.

### PROCEDURE:

1. Use a tissue or paper towel (not lens paper) to clean a slide and a coverslip. A microscope slide is a rectangular piece of glass about 2.5 cm x 7.5 cm. A coverslip is a thin, fragile piece of glass or plastic typically about 2.5 cm square.

2. Cut out a small letter "e" from the classified section of the newspaper. Place the "e" in the center of the glass slide.

3. Place a drop of water on the "e" using the medicine dropper and tap water.

4. Pick up the coverslip and touch one edge to the drop of water as shown in Figure 7. Slowly lower the coverslip to the slide. If this procedure is done correctly, you should get few or no air bubbles under the coverslip. If necessary, you can remove air bubbles by tapping the coverslip gently with the eraser end of a pencil. The slide will be studied in Part C.

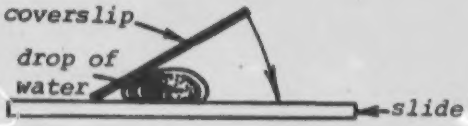

FIGURE 7: *Side view of slide.*

5. Prepare a slide (as in Steps 1 through 4) of a small piece of a colored picture from a newspaper or a magazine.

6. A third slide will now be made using an important technique called "teasing." If a thick specimen is placed on the stage, little or no light will pass through it, so it will be hard to see. Also, the many layers of cells will

make the picture confusing. So it is necessary to look at thin slices or sections.

In teasing, two dissecting needles are used to separate a piece of material into very small sections. Obtain a piece of cotton about 5 mm or less in diameter. Place the small piece of cotton in the center of the slide. Add two drops of tap water. Use the dissecting needles (one in each hand) to start separating the individual cotton fibers. Do this until you have the cotton fibers spread out over an area no larger than the area of the coverslip. Carefully place the coverslip on top. If you notice more than a few air bubbles, remove the coverslip, relocate the fibers and try again.

## C:. USE OF THE MICROSCOPE

### INTRODUCTION:

In this part of the activity, you will view microscopically the three slides you made in Part B. Be sure you are familiar with the terms and precautions discussed in Part A before continuing.

### PROCEDURE:

1. Turn the coarse adjustment so that the objectives are away from the stage.

2. It may be necessary to clean the objectives and eyepiece. Note: this should be done only with lens paper. The lenses could be damaged by tissue paper. Also, the lens paper should be folded twice so grease or sweat from the fingers will not soak through the paper and come in contact with the lenses.

3. Before you look through the eyepiece, try to get comfortable--otherwise microscopy can be very fatiguing. You should be seated in a way that will permit you to use the microscope without excessive stretching or hunching over. Some laboratory stools are adjustable to different heights. If necessary, exchange your chair or stool with another student.

4. Rotate the nosepiece so that the low-power objective is in position. If the microscope has a mirror, adjust the position of the mirror so that you can see a strong, even light in the eyepiece.

5. Place the slide of the "e" on the stage. Secure one end of the slide with a clip. Position the slide so that the "e" is in the center of the hole in the stage.

While watching the slide from the side, slowly turn the coarse adjustment. Lower the objective until it is close to the slide without touching it.

6. Under low power, the magnification is actually 10 (because of the eyepiece) x 10 (because of the objective), or 100 times. How much magnification can you get with the high-power objective?

7. Look through the eyepiece and slowly turn the coarse adjustment (away from the slide) until the "e" comes approximately into focus. If you pass the point of focus, look again from the side and lower the objective. REMEMBER: never move the objective toward the stage while looking through the eye piece! You might run the objective into the slide causing damage to the coverslip, slide and, worst of all, to the objective.

8. When the "e" is in approximate focus, turn the fine adjustment to sharpen the focus. Is the letter right side up or upside down? The "e" should appear upside down.

9. The "e" may not be in the center of the field of view. Move the slide to the left. Which way does the "e" move? Move the slide up. Which way does the "e" move? Now move the slide so as to center the "e".

10. While looking at the "e", move the lever or dial to adjust the lighting. Determine which positions of the lever or dial give the most light and which give the least light. Adjust the lighting so that it is neither too bright (glary) nor too dark.

11. Switch from low power to high power by slowly turning the nosepiece until the high-power objective is in place. Remember to watch from the side to be sure that the objective does not touch the coverslip. It should come very close without touching the coverslip. It may be necessary to raise the objective a little bit by turning the coarse adjustment knob before rotating the nosepiece.

12. Look through the eyepiece and turn ONLY the FINE ADJUSTMENT to focus. When using the high-power objective, never use the coarse adjustment to focus. You may have noticed that changing from low to high power did not entail a lot of focusing. Many microscopes are designed so that an object in focus at low power will also be in focus at high power.

13. How was the brightness changed when you switched from low to high power? If you're not sure, switch the objectives a few times. Make any necessary changes in the lighting.

14. Before removing the slide from the stage, rotate the nosepiece so that the low-power objective is in position and turn the coarse adjustment to move the objective away from the stage.

15. Repeat the procedure (Steps 5 through 14) to look at the other slides prepared in Part B--that is, the slide containing a piece of colored picture and the slide with the cotton fibers.

16. To see the effects of staining, focus the cotton fibers under low power. On one edge of the coverslip, place a drop of methylene blue stain. CAUTION: Methylene blue can stain your skin and clothing. Use a tissue or small piece of paper towel to absorb water from the other edge of the coverslip (see Figure 8). This process will cause the methylene blue to be drawn across the slide as the liquid is abosrbed from the other side. If the field is too dark, you can remove excess stain by adding water a drop at a time and absorbing it using the same technique.

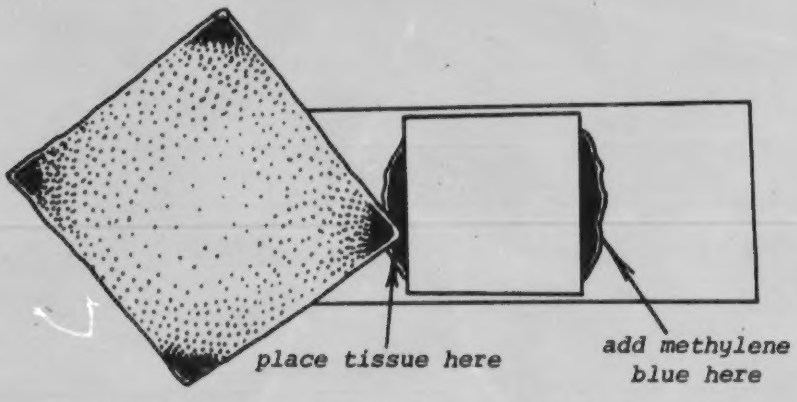

FIGURE 8: *Drawing stain across the coverslip.*

17. Remember to adjust the lighting as you make your observations.

18. When you are finished, be sure to wipe the stage dry of any water, and wash and dry the slides and coverslips. Also be sure that the low-power objective is in position before putting the microscope away. This way there is no possibility of the high-power objective coming into contact with the stage.

## D: MICROSCOPIC EXAMINATION OF TISSUES

### INTRODUCTION:

In this final part of the activity, you will have an opportunity to apply techniques of microscopy to the tissue samples collected in Part II. It is important to concentrate on good technique in your early exposure to microscopy because your approach to the microscope will soon become a habit. Good habits will improve your chances for success in many microscopic studies.

PROCEDURE:

1. Obtain the sections of tissue collected in Part II. A wet mount of each sample will be made,

a. Use a scalpel or razor blade to slice a very thin piece of cartilage off the ring. The piece does not need to be larger than 0.5 cm square but it should be as thin as possible. Why is this desirable? Make a wet mount of the cartilage tissue.

b. Slice a very thin piece of fat. After placing the coverslip on the slide, press with the eraser end of a pencil to flatten the fat into an even thinner layer.

c. Slice a thin, small section of lung tissue. This is difficult to do since the tissue is touch and elastic. Frozen lung tissue may be easier to slice.

d. Slice a thin section of epithelium. The epithelium should be aligned as in Figure 9 prior to slicing.

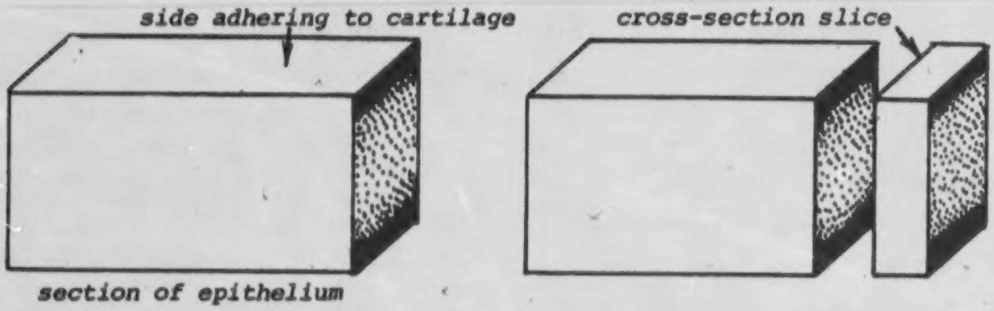

FIGURE 9: *Slicing a section of epithelium (schematic representation).*

e. Cut off a small section of muscle. Use two dissecting needles to tease the tissue apart. After teasing, you may need to discard some of the tissue before adding the coverslip. All of the tissue should fit under the coverslip.

2. Observe each of the tissues using the low-power objective (10X). Make a labeled sketch of each and describe what you see on your data sheet.

a. In the cartilage you should see small brown areas of cells on a white background. The background contains a flexible gelatin-like substance.

b. The fat should appear as a mass of small dark rings. The rings are actually spherical-shaped fat cells. Notice the three-dimensional appearance and the overlapping of the spheres.

c. The lung tissue should appear as a mass of tissue with many small bubbles. These are air bubbles within the alveoli. Remove the slide from the stage and press gently on the coverslip with the eraser end of a pencil. The pressure should squeeze the air bubbles out of the tissue. Look at the lung tissue through the microscope again.

d. The epithelium should have two different layers of cells.

e. The muscle tissue should include many fibers. The thickness of the fibers will depend on how much teasing you did. This kind of muscle is made up of hundreds of these long, thin cells which can contract and expand.

3. The tissues may be stained by adding a drop of methylene blue at the edge of the coverslip. The stain may be brought into contact with the tissues as in Part C, Step 16.

4. Stain and observe each of the tissues. What is the effect of staining on the appearance of the cells? Allow a few minutes for the stain to be absorbed in the cells. It may take some trial and error to get the best staining results. If you have time, experiment with the amount of stain used.

Remember to refocus and adjust the lighting as needed.

## LABORATORY ACTIVITY 13:

## PRESSURE

### INTRODUCTION:

Pressure is the ratio of the force exerted to the area over which the force is distributed. There are many examples of force. Whenever you push on something, you exert a force on it. Gravity exerts a force on you, and you exert this force on whatever you happen to be resting on. This force is called weight.

The relation between force and pressure is straightforward. The more force that is exerted, the greater the pressure. But the effect of area on pressure is somewhat more subtle. A proper understanding of the concept of pressure is essential to an understanding of most areas of biomedicine. The purpose of this activity is to help you to develop a feeling for what pressure is and how it is related to force and area.

In this activity, you will determine the pressure exerted on the bottoms of your feet under a variety of circumstances. In each case, the force exerted is your body weight. But the area over which that force is exerted differs, depending on whether you are standing on two feet or one, on tiptoe or flat-footed, barefoot or with your shoes on.

### MATERIALS:

bathroom scale
tray or pan
food coloring
drying oven
scissors
balance
4 sheets of paper

### PROCEDURE:

1. Remove your shoes and use the scale to determine your weight. Record the result. (This step may be done at any time during the period when the scale is free.)

2. Fill a tray or pan with tap water to a depth of 2 to 3 mm. Add 2 to 3 drops of food coloring and mix. Place a sheet of paper on the floor near the pan.

3. Remove your shoes and at least one sock. Dip your bare foot into the colored water, shake off the excess water and then stand on the sheet of paper. Be sure you put your foot down so that all of it is on the paper.

4. Blot the excess water from the paper, and set it aside to dry.

5. Repeat Steps 3 and 4, but standing on tiptoe this time.

6. Wash off the coloring and dry your foot.

7. Initial your two sheets of paper (not on the footprint). Then place them in the oven to finish drying.

8. On a third sheet of paper, draw an outline of the part of your foot that is supported by your shoe when you are standing in it. Start by drawing an outline of your bare foot. (You should be able to do this by yourself, while sitting in a chair.) The difficult part of this step is in knowing where to draw the line in the region of the instep. Check the width of the inner sole of your shoe in the instep region. Then correct your drawing so that the width of the outline is approximately the same as that of the inner sole.

9. Put your sock and shoes back on.

10. Cut out the foot outline you made in Step 8.

11. Determine and record the mass (± 0.01 g) of a complete sheet of paper and of the cut-out of your foot.

12. When your other two sheets of paper are dry, remove them from the oven. Cut out the colored footprint and toe-print. Determine and record the masses of these two cut-outs.

13. Use the following information to calculate and record your body weight in grams. 1 lb ≈ 453.6 g. However, to keep the calculations simple, you may use the following approximation instead. 1 lb ≈ 450 g. The error introduced in this way will be quite small.

$$\text{weight } \cancel{\text{in lb}} \times \frac{450\text{ g}}{\cancel{1\text{ lb}}} \approx \text{weight in g}$$

14. Use the following information to calculate and record the area to the nearest sq cm of each of your cut-outs. A standard 8.5 x 11 inch sheet of paper has an area of 93.5 sq in. This is equivalent to approximately 603 sq cm. The area of a cut-out may be found from the following proportionality.

$$\frac{\text{area of cut-out}}{\text{area of sheet}} = \frac{\text{mass of cut-out}}{\text{mass of sheet}}$$

$$\text{area of cut-out} = \frac{(\text{area of sheet})(\text{mass of cut-out})}{\text{mass of sheet}}$$

$$\text{area of cut-out} \approx \frac{603(\text{mass of cut-out})}{\text{mass of sheet}}$$

15. Calculate and record the pressure exerted on the bottoms of your feet, toes, etc., under the following conditions.

a. When you are standing barefoot on two feet. (Remember to double the area of your cut-out.)

b. When you are standing in your shoes on both feet.

c. When you are standing barefoot on tiptoe on both feet.

d. When you are standing barefoot on tiptoe on one foot.

## DISCUSSION QUESTIONS:

1. How does the pressure exerted while standing on two feet compare with the pressure exerted while standing on one foot?

2. Compare your results with those of others. Can you find persons who weigh more than you do, but exert less pressure on their feet? Can you find opposite examples, i.e., persons who weigh less than you, but exert more pressure on their feet? How can such results be accounted for?

3. Do our bodies sense force or pressure? (Think about squeezing a thumbtack between thumb and forefinger.)

4. A shoe that is properly designed and fits properly should provide support for almost all of the bottom of the foot. Compare walking with and without shoes in terms of the pressures exerted on the feet.

5. Which of your calculated pressures is the greatest? Why?

6. Which exerts less pressure on the ground: a 700-lb gorilla on roller skates, or an 800-lb gorilla on identical roller skates?

7. You want to cross an ice-covered pond in wintertime. Why is the ice less likely to crack if you lie down and crawl, rather than walk across?

8. Discuss the respective advantages and disadvantages of boots, skis, snowshoes and riding horseback in crossing deep, soft snow.

## SAMPLE DATA SHEET:

body weight in lb:

mass of full sheet of paper:

mass of cut-out (for area in contact with shoe):

mass of cut-out (barefoot):

mass of cut-out (barefoot, on tiptoe):

calculated body weight in g:

calculated area of cut-out (for area in contact with shoe):

calculated area of cut-out (barefoot):

calculated area of cut-out (barefoot, on tiptoe):

calculated pressure, barefoot on two feet:

calculated pressure, standing in shoes, both feet:

calculated pressure, barefoot on tiptoe, both feet:

calculated pressure, barefoot on tiptoe, one foot:

## LABORATORY ACTIVITY 14:

## DETERMINATION OF AEROBIC CAPACITY

### INTRODUCTION:

Exercise increases the body's energy requirements. This increased demand for energy is satisfied through chemical processes in the body that use oxygen. Fatigue results when the supply of oxygen cannot meet the demands of the body. Physical fitness is a measure of the body's ability to meet its oxygen requirements during prolonged exercise.

A commonly used measure of physical fitness is one's aerobic capacity per unit of body weight. Aerobic capacity refers to the amount of oxygen that the body uses during maximum exertion. This measure ranges between 30 and 50 ml of oxygen per minute per kg of body weight (ml/kg-min) for untrained persons. For trained athletes, the aerobic capacity may exceed 80 ml/kg-min.

It is possible to measure aerobic capacity without undergoing prolonged physical exercise. A five-minute "step test" is performed that increases the pulse rate significantly, but does not tax the body to its limit. The aerobic capacity is determined from the final pulse rate and amount of work done during the test. A mathematical device called a "nomogram" is used to simplify the calculations.

The step test and the associated nomogram were developed from measurements on reasonably active young adults. It is possible for a person in fair-to-poor condition to develop a pulse rate that is too high for the nomogram. Since such a result could be unsafe, it is recommended that persons who are very inactive, excessively overweight, or who smoke regularly, perform the step test at half the normal rate. By reducing the amount of exercise, the pulse rate is lowered, and valid results can be obtained.

### MATERIALS:

bathroom scale
clock with second hand
source of timed signals
step, 30 to 35 cm high
meter stick

### PROCEDURE:

1. Weigh yourself and record your weight on Data Sheet 14. If you intend to perform the step test without shoes, remove them before weighing.

2. Measure and record the height of the step that is being used.

3. Pair up with another student. Practice taking each other's pulse for 15-second intervals. Repeat this several times until it can be accomplished without difficulty. The pulse is taken by placing the index and middle finger of one hand on the inner side of your partner's wrist. The pulse may be felt in the radial artery which is located about 3 cm above the wrist joint, near the base of the thumb. Since a 15-second pulse count may be off by one beat in either direction, multiplying the result by four to obtain the number of beats per minute will have an imprecision of ±4 beats.

4. The instructor will provide a source of timed signals that is set to produce 45 signals per minute. If you step up or down at each signal, the step rate will be $22\frac{1}{2}$ per minute (the normal step rate). If you step up or down at every other signal, the step rate will be $11\frac{1}{4}$ per minute and the amount of work done half as much as at the normal step rate. Select a step rate of either $22\frac{1}{2}$ or $11\frac{1}{4}$ steps per minute and record the rate on your data sheet. (The introduction explains how to choose between these two rates.)

5. Practice a few steps to get the proper rhythm. The exercise is performed by placing one foot on the step, stepping up to an erect position with the weight on both feet, and stepping down with the same foot used in stepping up.

6. Perform the step test for approximately 5 minutes. The exact length of time is not critical, since the pulse should reach a steady rate before the end of the test. Follow the timer carefully, but do not be concerned if a beat is occasionally missed.

7. After 5 minutes, have your partner measure your 15-second pulse rate. Record the result. Since your pulse begins to slow down after the exercise is completed, be sure that it is taken immediately after stopping.

8. Convert your pulse rate to beats per minute and indicate the imprecision (±4).

9. If you stepped at the lower rate and your final one-minute pulse measures less than 122 for a male or 120 for a female, repeat the exercise at the higher rate.

10. If you stepped at the higher rate and your one-minute pulse measures more than 170 for a male or 172 for a female, relax for a while. Then repeat the exercise at the lower rate, if time and energy permit.

11. Calculate and record your weight in kg. (See the sample calculation sheet, item 1.)

12. Calculate and record your work rate. (Sample calculation sheet, item 2.)

13. Use the nomogram on the following page to determine and record your aerobic capacity. To read the nomogram, run a straightedge from the midpoint of your one-minute pulse rate on the left to your work rate on the right scale. The point where the line crosses the center scale gives your aerobic capacity. (Sample calculation sheet, item 3.)

14. Repeat Step 13 for a pulse rate 4 beats per minute greater than the midpoint and four beats per minute less than the midpoint.

15. Calculate and record your fitness index for each of the three recorded aerobic capacities. (Sample calculation sheet, item 4.)

16. Determine your fitness category according to the following table.*

| FITNESS CATEGORY | FITNESS INDEX in ml/kg-min |
|---|---|
| Very Poor | <25.0 |
| Poor | 25.0 - 33.7 |
| Fair | 33.8 - 42.5 |
| Good | 42.6 - 51.5 |
| Excellent | >51.5 |

SAMPLE DATA TABLE:

(Note that sample data have been included. These data are used in the sample calculations that follow.)

Body weight: 218 lb
Step height: 34.6 cm
Step rate: 22.5 steps per minute
15-second pulse rate: 34 ± 1 beats per 15 sec
one-minute pulse rate: 136 ± 4 beats per min

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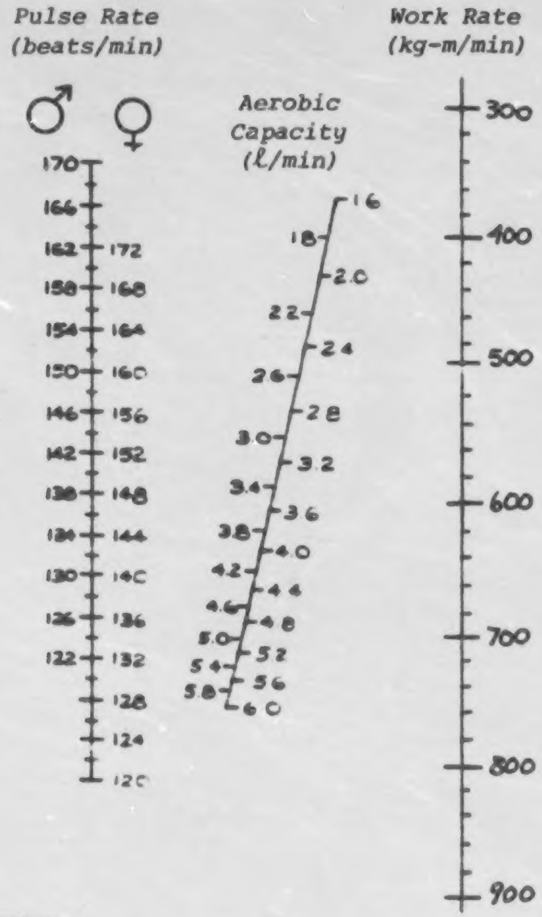
*From Cooper, Kenneth H. The New Aerobics. Bantam Books (#553-04667), New York, 1970, p. 28.

AEROBIC CAPACITY NOMOGRAM*



SAMPLE CALCULATION SHEET:

1. Body weight in kg:

$$218 \text{ lb} \times \frac{1 \text{ kg}}{2.2 \text{ lb}} = 99.1 \text{ kg}$$

2. Work rate in kg-cm per min = (weight in kg)(step height in cm)(steps/min)

$$= (99.1)(34.6)(22.5)$$

$$= 77{,}149 \text{ kg-cm/min}$$

Work rate in kg-m per min:

$$\frac{77{,}149 \text{ kg-cm}}{\text{min}} \times \frac{1 \text{ m}}{100 \text{ cm}} = 770 \text{ kg-m/min}$$

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*From Astrand, Irma. Aerobic Work Capacity in Men and Women With Special Reference to Age. Acta Physiologica Scandinavica, Vol. 49, Supp. 169, 1960, p. 51.

3. From nomogram (assuming subject is a male):

aerobic capacity $\approx$ 4.6 liters/min, for pulse rate of 136 beats/min

$\approx$ 4.4 liters/min, for pulse rate of 140 beats/min

$\approx$ 4.9 liters/min, for pulse rate of 132 beats/min

4. Aerobic capacity in ml/min:

$$\frac{4.6 \text{ liters}}{\text{min}} \times \frac{1000 \text{ ml}}{1 \text{ liter}} = 4600 \text{ ml/min}$$

The other two pulse rates give 4400 and 4900 ml/min.

$$\text{Fitness index in ml/kg-min} = \frac{\text{aerobic capacity in ml/min}}{\text{weight in kg}}$$

$$= \frac{4600}{99.1} \approx 46.4 \text{ ml/kg/min}$$

The other two pulse rates give

$$\frac{4400}{99.1} \approx 44.4 \text{ ml/kg-min} \qquad \text{and} \qquad \frac{4900}{99.1} \approx 49.4 \text{ ml/kg-min}$$

All three pulse rates indicate a fitness category of "good."

DISCUSSION QUESTIONS:

1. To what extent was your fitness index affected by the imprecision of the pulse rate measurement? Could a change of just one beat in the 15-second pulse count change the final fitness category? Did it in your case?

2. Compare your results with those of other students. Are there any smokers in the class? How did their results compare with those of non-smokers? How did the athletes compare with the rest of the class?

3. If your fitness category is below "excellent," how do you think you could improve it?

## LABORATORY ACTIVITY 15:

## THE EFFECT OF TEMPERATURE CHANGE ON THE VOLUME OF A CONFINED GAS

### INTRODUCTION:

From everyday experience, you may already know that the volume of a gas is affected by a change in temperature. For example, we are warned that aerosol cans may explode if they are thrown into a fire. As the temperature rises, the gas inside expands until the container bursts with explosive force.

Tho purpose of this activity is to collect numerical data on the relation between the temperature and the volume of a gas. From the data collected, a mathematical expression can be developed that will describe the general relationship between these two variables. Since pressure also has an effect on the volume of a gas, the pressure will be held constant during the investigation.

In order to measure the volume of a sample of gas, it is necessary to confine the gas within a container. In addition, the container must be able to expand and contract as the volume of gas changes. In this activity a sample of air will be trapped in a section of glass tubing. The tube will be sealed at the top. The bottom of the air sample will be in contact with the surface of a liquid. The level of the surface of the liquid will rise and fall in response to contraction and expansion of the air sample. (See Figure 1.)

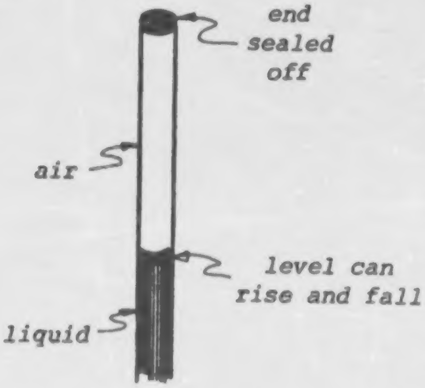

FIGURE 1: *Confinement of the air sample.*

The liquid used to confine the gas sample will be ethylene glycol, which is commonly known as "anti-freeze" because of its use in car radiators. You might wonder why water is not used instead. The reason for this is that water vapor rising from the surface of the water would increase the volume of confined gas and produce misleading results. At the temperatures to be used in the investigation, the amount of ethylene glycol vapor produced is small enough to be ignored.

Since the volume of air cannot be measured directly, the height of the air column inside the glass tubing will be measured instead. The actual volume of gas is given by the formula for the volume of a cylinder

$$V = \pi r^2 h$$

where $\pi \approx 3.14$, $r$ is the inside radius of the glass tubing and $h$ is the measured height of the air column. Since $\pi$ and $r$ are constant for all measurements, the volume is proportional to the height.

## MATERIALS:

- 2 lengths glass tubing, 30 cm long
- length glass tubing, 5 cm long
- 2 lengths rubber tubing, 25 cm long
- rubber stopper, #4, two-hole
- silicone rubber cement
- ethylene glycol
- pinch clamp
- clear tubing (glass or plastic), 22 mm I.D., 30 cm long
- thermometer, 0 - 100° C
- ring stand
- 3 ring-stand clamps
- millimeter ruler
- 2 beakers, 250-ml
- medicine dropper
- ice
- glycerine
- tape

## PROCEDURE:

1. Insert one of the 30-cm lengths of glass tubing through the rubber stopper as shown in Figure 2. Note that the wide end of the stopper is down.

2. Complete the initial assembly by connecting the two 30-cm lengths of glass tubing with a length of rubber tubing. (See Figure 2.)

3. Obtain about 7 ml of ethylene glycol in a beaker.

4. Have a second person help you with this step. Use a medicine dropper to add ethylene glycol to the tube marked B in Figure 2. Continue adding

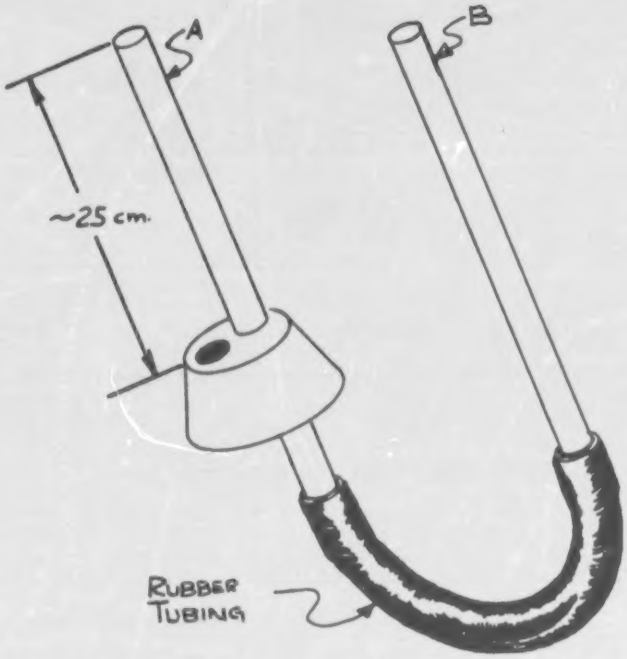

FIGURE 2: *Initial assembly.*

liquid until the meniscus is approximately 20 cm from the top of each tube. Do not allow the liquid to come in contact with the end of Tube A.

5. Dislodge any trapped air bubbles by tapping the tubing.

6. While holding both tubes in a vertical position, squeeze a small amount of silicone rubber cement into the end of Tube A. The cement plug should be 1 to 2 cm long.

7. Store the initial assembly overnight to allow the cement to harden. Storage should be in a roughly vertical position, so that the liquid cannot reach either the cement in Tube A or the open end of Tube B. (It may also be a good idea to cover the open end of Tube B to prevent dust particles from entering.)

8. When the cement plug has hardened, complete the assembly of the apparatus as detailed in the following steps. (See Figure 3.)

9. Clamp the large glass (or plastic) tube to the ring stand. This tube is called the "jacket tube" in Figure 3.

10. Insert a short piece of glass tubing into the empty hole in the stopper. Attach to this tube a length of rubber tubing and a pinch clamp.

11. Insert the stopper into the lower end of the jacket tube. Adjust the stopper so that the tube containing the air sample does not touch the inner wall of the jacket tube. Seat the stopper securely so that it does not fall out later during the collection of data.

12. Clamp the leveling tube to the ring stand.

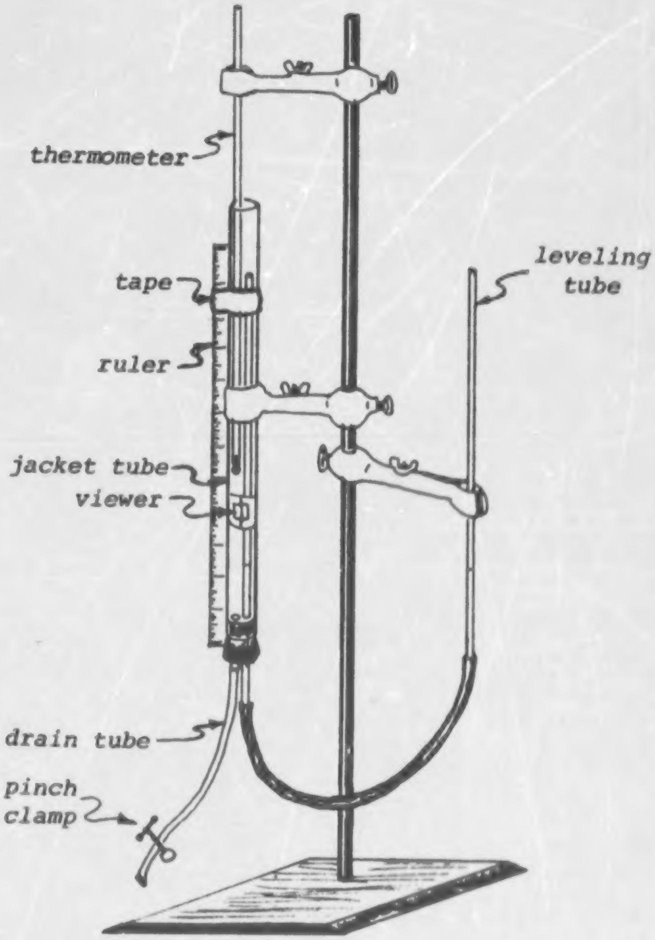

FIGURE 3: *Final assembly.*

13. Attach the thermometer as shown in Figure 3. The clamp should be above the 60-degree mark on the thermometer. The bulb of the thermometer should be approximately halfway down the length of the air column. The thermometer may touch the tube containing the air sample, but it should not touch the jacket tube.

14. In order to obtain accurate measurements of the length of the air column, it is important to keep the eye at the same level as the meniscus. An easy way to do this is to take a strip of paper approximately 2 cm by 10 cm, wrap it into a cylinder around the jacket tube, and secure it with tape. When this "viewer" is slid into the proper position, the level of the meniscus can be read easily. (See Figure 4.)

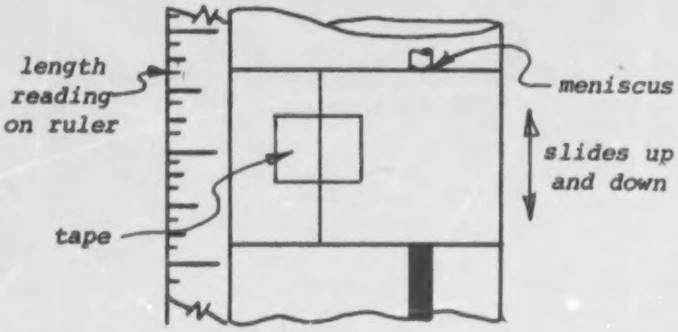

FIGURE 4: *The viewer.*

15. Tape a millimeter ruler to the outside of the jacket tube. The zero mark on the ruler should line up with the lower edge of the cement plug, so that the length of the column of air can be read directly off the ruler. The upper piece of tape should be below the zero-degree mark on the thermometer so that it will not obscure any temperature readings. The lower piece of tape (not shown in Figure 3) should be wrapped around the lowest part of the jacket tube so that it will not interfere with the reading of the meniscus.

16. If the end of the leveling tube was covered for storage purposes, remove the cover now.

17. Before beginning to collect data, it is important to understand how the apparatus works. Hot or cold water poured into the jacket tube changes the temperature of the trapped air, which can be read with the thermometer. The change in temperature causes the gas to expand or contract, which in turn changes the level of the ethylene glycol meniscus. In order to keep the trapped gas at a constant pressure, the leveling tube is adjusted so that the ethylene glycol is kept at the same level in both tubes. If this is done, the trapped gas will be at atmospheric pressure for each set of readings.

18. Fill the jacket tube with hot tap water.

19. While the thermometer is adjusting to the new temperature, move the leveling tube up or down until its meniscus is at the same height as the meniscus inside the jacket tube.

20. Adjust the viewer in order to read the level of the meniscus below the air column.

21. Record the temperature to the nearest 0.5 °C and the length of the air column to the nearest 0.05 cm.

22. Repeat steps 18 through 20 at successively lower temperatures until a temperature near 0 °C is reached. Eight to ten sets of readings in all should be sufficient. At first, while the hot water is in the jacket tube, it may be simplest to allow the water to cool by itself, taking a set of readings every 5 °C or so. When this procedure becomes too time-consuming, empty the jacket tube through the drain tube and re-fill it with combinations of hot and cold tap water mixed in various proportions. Use combinations of cold water and ice water to obtain temperatures below that of cold tap water.

23. The collected data will be analyzed in the mathematics class.

SAMPLE DATA TABLE:

| TEMPERATURE (nearest 0.5 °C) | LENGTH (nearest 0.05 cm) |
|---|---|
| | |
| | |
| | |

DISCUSSION QUESTIONS:

1. Explain how the leveling tube permits one to maintain the pressure of the trapped air at atmospheric pressure.

2. Why shouldn't the thermometer have been in contact with the jacket tube?

3. What is the imprecision of your temperature measurements? Of your length measurements?

4. Calculate the volume of the air sample for the highest temperature and the lowest temperature you recorded. Use $\pi \approx 3.14$ and radius $\approx 0.2$ cm. What are the units for the calculated volumes?

## LABORATORY ACTIVITY 16:

## THE EFFECT OF PRESSURE CHANGE ON THE VOLUME OF A CONFINED GAS

### GENERAL INTRODUCTION:

As we have seen earlier, the process of breathing depends upon the interaction of the pressure and volume of the air inside our lungs. In fact, if pressure and volume were not related properties of a gas, we could not breathe at all. An understanding of this relationship is essential to an understanding of normal lung function as well as many disorders of the lung.

In 1662, an Englishman named Robert Boyle first described the relationship between the pressure and volume of a gas. The relationship he found is approximately true for all gases. In this activity you will collect data which may be used to determine the nature of this relationship, which has become known as Boyle's Law.

Since temperature also has an effect on the volume of a gas, it is necessary to hold temperature constant while investigating the relation between pressure and volume. In this activity, it is assumed that the temperature of the laboratory remains constant during the collection of data. Any slight changes that occur in the temperature of the laboratory should have a negligible effect upon the data.

Part I and Part II are two separate procedures for obtaining the data. Use the one for which equipment is available.

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### PART I:

### INTRODUCTION:

In this procedure a sample of air is trapped inside a syringe. Pressures are exerted on the air sample by placing weights on the plunger of the syringe. The amount of weight applied and the resulting volume of the air are recorded. The corresponding pressures are later calculated from the forces exerted by the weights and the cross-sectional area of the inside of the syringe.

### MATERIALS:

plastic syringe-platform instrument
set of identical weights
glycerine
balance, ± 0.01 g

### PROCEDURE:

1. When your instructor gives you your weights, he will tell you how much one of them weighs. Record this value.

2. Remove the syringe from its base, and remove the rubber cap from the tip of the syringe. Pull the plunger out of the barrel of the syringe. Use a small amount of glycerine to lubricate the inside of the barrel and the rubber part of the plunger. Move the plunger in and out of the barrel until it moves smoothly.

3. While holding the syringe in a horizontal position, move the plunger until the rubber seal rests on the 35-cc (milliliters) mark. Place the rubber cap onto the tip of the syringe. Then place the syringe into the base. This forms an airtight seal so that gas can neither enter nor escape from the syringe. Record the initial volume of 35.0 ml with no weight on the confined gas.

4. Place enough weights on the platform so that the total weight on it is between 700 and 1000 grams. Record the number of weights.

5. Pull the platform upward slightly and let go of it. Read the position of the rubber seal to the nearest 0.1 ml, and record it. Push down on the platform and then remove your hand. Record the new position to the nearest 0.1 ml.

By taking two comparable readings each time you add weights, you will be able to establish a range of imprecision for each volume you obtain. Can you explain why the "from above" and "from below" values are different? (1)

6. Repeat Steps 4 and 5 using an additional 700 to 1000 grams of weights. Remember to enter in the "number of weights" column the total number of weights on the platform, and not the number of weights added. Continue this procedure until you have run out of weights.

7. Disassemble the apparatus and determine the weight of the platform and plunger to the nearest gram. Record this value.

SAMPLE DATA SHEET:

Initial volume with no weight on the confined gas: 35.0 ml

| Number of weights | Volume of air in syringe (ml) |
|---|---|
| | from above |
| | from below |
| | from above |
| | from below |

Weight of a single weight: ________ grams

Weight of platform and plunger: ________ grams

CALCULATIONS:

1. Use the following information to convert the value in the "number of weights" column into pressure expressed in grams per square centimeter. Show all calculations.

If P = the pressure exerted on the confined gas, then

$$P = \frac{F}{A}$$

where F = the force in grams exerted by the plunger, platform and weights and A = the area over which force F is exerted, which is the cross-sectional area of the inside of the barrel of the syringe.

$$A = 4.4 \text{ sq cm}$$

This value was calculated from one measurement taken along the side of the barrel of the syringe. Can you describe the measurement and the calculation? (2)

F = (no. of weights)(wt. of one weight) + wt of plunger and platform

$$P = \frac{F}{A} = \frac{\text{(no. of wts.)(wt. of 1 wt.) + wt. of plunger and platform}}{4.4}$$

EXAMPLE:

No. of weights = 4

Weight of one weight = 416 grams

Weight of plunger and platform = 173 grams

$$P = \frac{4(416) + 173}{4.4}$$

$$= \frac{1837}{4.4}$$

$$= 418 \text{ g/sq cm}$$

2. The true value of the volume for each pressure is assumed to lie somewhere between the two values recorded. Use the following information to calculate the range of imprecision for each recorded set of "from above" and "from below" volumes. Show all calculations.

If $V_a$ = volume "from above", and $V_b$ = volume "from below", then the midpoint is the average of the two readings, or

$$\frac{V_a + V_b}{2}$$

The imprecision is

$$\pm \frac{V_a - V_b}{2}$$

Thus the range of imprecision is

$$\frac{V_a + V_b}{2} \pm \frac{V_a - V_b}{2}$$

EXAMPLE:

Volume from above = 20.2 ml

Volume from below = 18.0 ml

$$\text{Range of imprecision} = \frac{20.2 + 18.0}{2} \pm \frac{20.2 - 18.0}{2}$$

$$= \frac{38.2}{2} \pm \frac{2.2}{2}$$

$$= 19.1 \pm 1.1 \text{ ml}$$

3. Complete a new table of the calculated values as follows. Note that the first data pair are a special case. With no weights on the confined gas, the pressure was zero. Since the initial volume was set at 35.0 ml, it is not shown as a range of imprecision.

| Pressure exerted in g/sq cm | Range of imprecision of volume in ml |
|---|---|
| 0 | 35.0 |
| | |
| | |

4. Graph your data, using pressure as the horizontal coordinates and volume as the vertical coordinates. Assuming that your smallest volume is greater than 14 ml, ΔV will be 35 - 14 = 21. If we use volume on the long axis, we have $\frac{21}{24}$ for the big units, which is almost 1. Using 1 ml for the big units will make each small unit .2 ml. Assuming that your ΔP may be as large as 1400, we have on the short axis, $\frac{1400}{18}$ for the big units, which is almost 78. Using 100 g/sq cm for the big units will make each small unit 20 g/sq cm.

## PART II:

## INTRODUCTION:

In this procedure a sample of air is trapped inside a length of glass tubing, much as in Laboratory Activity 15. The glass tube is plugged at the top and the air inside is in contact with the surface of a liquid at the bottom. Pressure is applied to the air sample by the weight of a column of liquid, which is just the reverse of the way a manometer works. The pressure is varied by changing the height of the column of liquid.

The height of the liquid and the resulting length of the air column are recorded. The corresponding pressures are later calculated from the recorded heights and the density of the liquid.

The liquid used in the activity is sodium bromide dissolved in water. The reason for using this liquid, rather than pure water, is that it has a density of 1.5 g/cu cm, which will permit the exertion of greater pressures than would be possible with pure water.

## MATERIALS:

40 ml sodium bromide fluid
length rubber tubing, 150 cm long
2 lengths glass tubing, 30 and 35 cm long
ring stand and clamp
beaker, 50-ml
pipet, 10-ml
millimeter ruler
silicone rubber cement
meter stick
wooden dowel or stick, at least 120 cm long
tape

## PROCEDURE:

1. Obtain approximately 40 ml of the sodium bromide fluid in a 50-ml beaker.

2. Begin assembly of the apparatus by attaching a piece of glass tubing to each end of the rubber tubing. The longer glass tube will contain the trapped air sample and will be called the air tube. The shorter glass tube will be used to change the pressure on the trapped air and will be called the pressure tube. See Figure 1 on the following page.

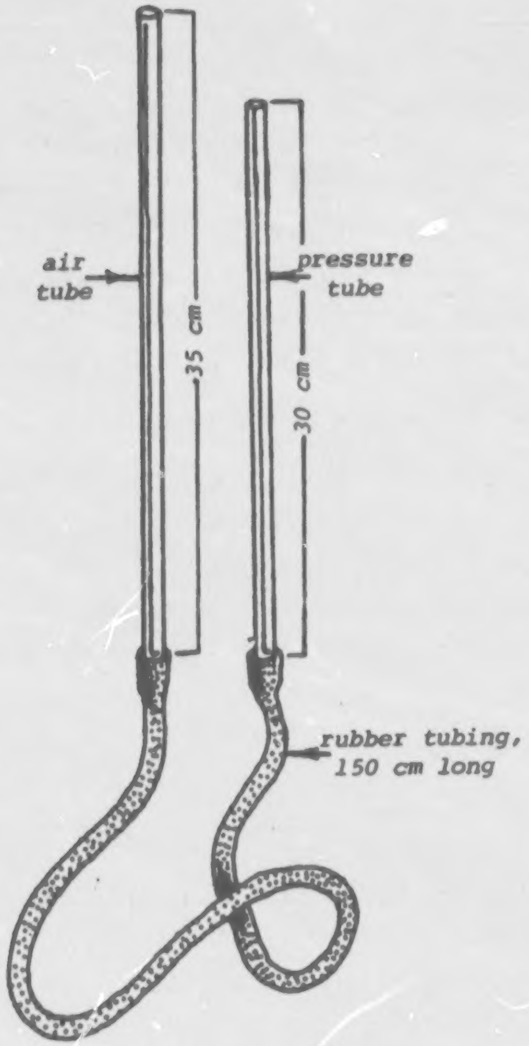

FIGURE 1: *Initial assembly of the apparatus.*

3. Use a 10-ml pipet to begin filling the pressure tube with the sodium bromide fluid. Do not get the end of the air tube wet.

4. With the two tubes held in the position shown in Figure 1, continue adding fluid to the pressure tube until the fluid level is approximately 20 cm from the top of the air tube.

5. Squeeze a small plug of silicone rubber cement, about 1 cm long, into the end of the air tube.

6. Store the apparatus overnight to allow the cement to harden. Storage should be in a roughly vertical position, so that the fluid cannot reach either the cement in the air tube or the open end of the pressure tube. (It may also

be worthwhile to cover the end of the pressure tube to prevent dust particles from entering.)

7. When the cement had hardened, complete the assembly of the apparatus as shown in Figure 2 (on the following page). The ruler should be positioned so that the zero point coincides with the top of the air column (i.e., the bottom of the cement plug).

8. The wooden dowel is used to measure height in this procedure. With the help of a meter stick, mark off 20-cm intervals on the dowel. It may help to label the marks "0", "20", "40", etc.

9. To take the readings, the bottom of the dowel is held even with the meniscus in the air tube, while the meniscus in the pressure tube is moved to the proper mark on the dowel. Record the height difference between the two fluid levels.

10. While maintaining the height difference just recorded, determine and record the length of the air column to the nearest 0.1 cm. Remember, in all readings, to keep the eye at the same level as the meniscus.

11. Vary the height difference between the two fluid levels in 20-cm intervals, starting near the ceiling and working downward. Readings in which the pressure tube is below the air tube should be recorded as negative.

SAMPLE DATA TABLE:

| Height difference of fluid (cm) | Length of air column (nearest 0.1 cm) |
|---|---|
| 120 | |
| 100 | |
| 80 | |
| -80 | |
| -100 | |
| -120 | |

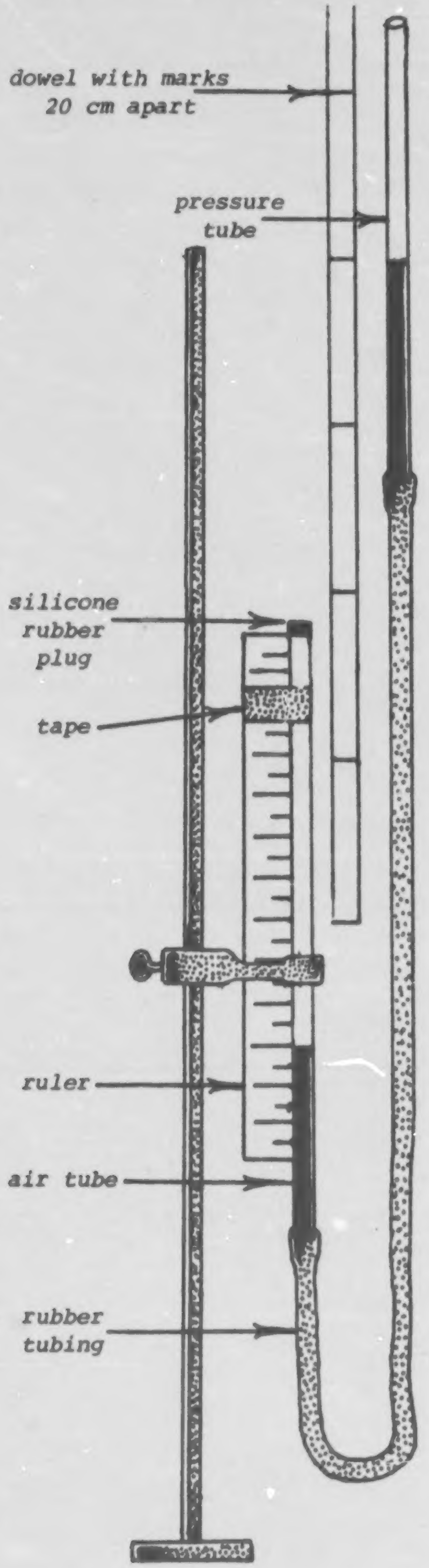

FIGURE 2: *Final assembly of the apparatus.*

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CALCULATIONS:

1. As was discussed in Section 13, the pressure exerted by a column of fluid can be found by multiplying the density by the height of the column. The density of the sodium bromide fluid is 1.5 g/cu cm. Therefore the pressure in g/sq cm can be found for each height difference in your data table by multiplying the height difference by 1.5.

Since the volume of trapped air is proportional to the length of the air column, we can treat the length of glass tubing that was filled with air as if it were a unit of volume. "cm of tubing" is, of course, not a conventional volume unit, but it saves us the trouble of finding the radius of the air column and making lengthy calculations.

Complete a new table of data showing the calculated pressures and the same air-column lengths as before.

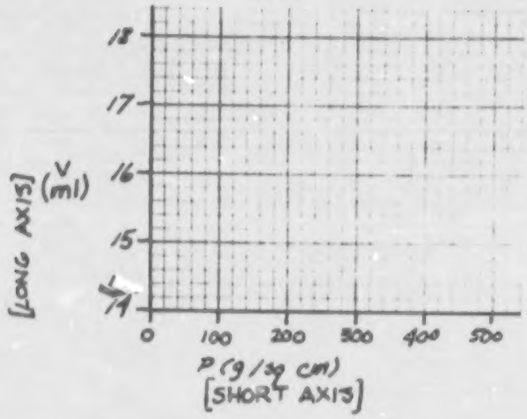
| Pressure (g/sq cm) | Volume (cm of tubing) |
|---|---|
| | |
| | |
| | |

2. Graph your data, using pressure as the horizontal coordinates and volume as the vertical coordinates. Assuming that your total range of volumes is no greater than 7.2 cm, $\Delta v$ will be 7.2. If we use volume on the long axis, we have $\frac{7.2}{24}$ or 0.3 for the big units. Using 0.3 cm for the big units will make each small unit 0.06 cm. Assuming that your $\Delta P$ may be as large as 360, we have on the short axis, $\frac{360}{18} = 20$ g/sq cm for the big units. This will make each small unit 4 g/sq cm.

3. When you have completed the graph, hand it in along with all data and calculations. Be sure your name is on all sheets. The materials will be returned to you in the mathematics class for further analysis.

4. Graph your data, using pressure as the horizontal coordinates and volume as the vertical coordinates. Assuming that your smallest volume is greater than 14 ml, $\Delta V$ will be $35 - 14 = 21$. If we use volume on the long axis, we have $\frac{21}{24}$ for the big units, which is almost 1. Using 1 ml for the big units will make each small unit .2 ml. Assuming that your $\Delta P$ may be as large as 1400, we have on the short axis, $\frac{1400}{18}$ for the big units, which is almost 78. Using

100 g/sq cm for the big units will make each small unit 20 g/sq cm.



Show each ordered pair except the first as a range of imprecision in the vertical dimension. Use a dot in the center to indicate the location of the midpoint (I).

5. When you have completed the graph, hand it in along with all data and calculations. Be sure your name is on all sheets. The materials will be returned to you in the mathematics class for further analysis.

# LABORATORY ACTIVITY 17:

# THE EFFECT OF EXTERNAL TEMPERATURE ON RESPIRATION

## INTRODUCTION:

In Section 17-3, we discussed the respiratory response of man and other homeothermic animals to cold. When the body is subjected to unusual heat losses, there is a speed-up in the chemical processes in the cells that produce heat. In this way, a constant internal temperature is maintained.

In most of these chemical processes, oxygen is used up and carbon dioxide is generated. When we are subjected to cold, our breathing rate automatically increases. We both consume oxygen and eliminate carbon dioxide at a greater rate.

In this activity you will have an opportunity to compare the difference in the breathing of a mouse at room temperature and at a temperature of about 10 °C. The instrument used in this investigation is called a respirometer. As shown in Figure 1 in the procedure, it consists of two closed chambers connected by a manometer tube. The chamber on the left contains the experimental mouse and a small bag of soda lime. The soda lime absorbs the carbon dioxide exhaled by the experimental mouse.

The chamber on the right side contains a "control" mouse, but no soda lime. The function of the control mouse is to have conditions on the two sides of the manometer as nearly identical as possible, except for the one factor being investigated. In this way, we can be reasonably sure that any changes in the manometer level are caused by the absorption of carbon dioxide by the soda lime.

As carbon dioxide is absorbed on the experimental side, the volume and pressure of the air on that side of the manometer is reduced. This causes the manometer level to rise on the experimental side. The change in the level of the manometer fluid provides a measure of the amount of carbon dioxide absorbed by the soda lime, and is an indirect measure of the breathing rate of the experimental mouse.

## MATERIALS:

2 mice
mouse food
soda lime
2 Erlenmeyer flasks, 250-ml, with 2-hole stoppers
beaker, 250-ml (or other size)
pipet, 10-ml
glass tubing, 1 30-cm length and 4 short lengths to fit stoppers
rubber tubing, 2 60-cm lengths, 1 30-cm length and 2 short lengths

water-bath container
2 ring stands
2 ring-stand clamps
2 screw clamps
thermometer
ice
cheesecloth
scissors
string
masking tape

PROCEDURE:

1. Connect the pipet and the long piece of glass tubing to the two ends of the 30-cm length of rubber tubing. Using mouth suction, fill the tubing with water so that when the tubes are placed side by side, the liquid level is near the 8-ml mark (see Figure 1). Attach the tubing with masking tape to the outside of the water bath container.

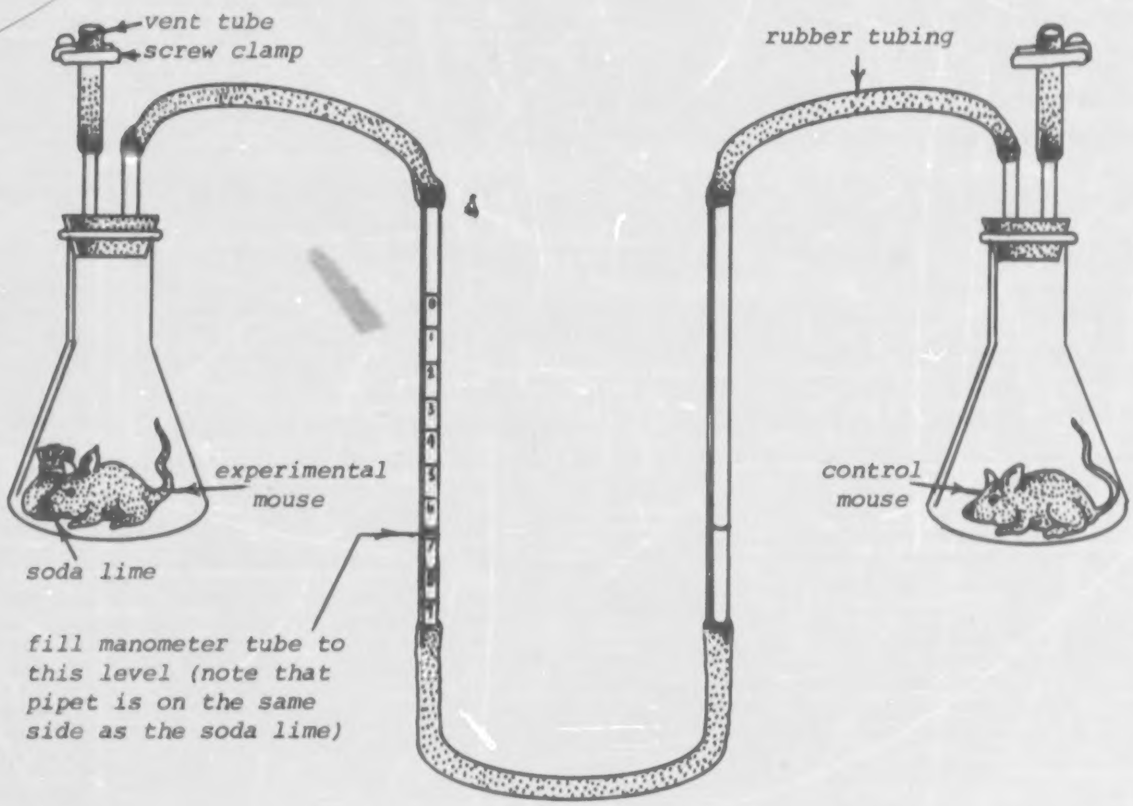
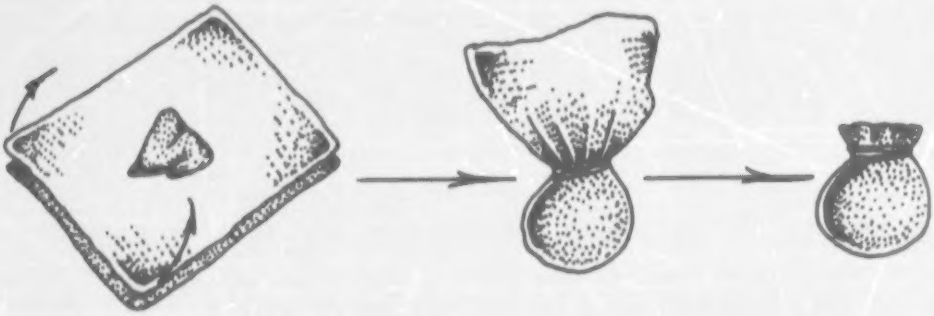

vent tube
screw clamp
rubber tubing
experimental mouse
control mouse
soda lime
fill manometer tube to this level (note that pipet is on the same side as the soda lime)

FIGURE 1: *Assembly of the respirometer.*

2. Complete the assembly of the apparatus as shown in Figure 1, and place the two flasks in the empty water-bath container. It is not necessary to add water to the water bath until Step 11.

3. Cut off a length of cheesecloth and fold it in half in order to obtain a double thickness. Place a small quantity of soda lime in the center of the cheesecloth, gather the corners together and tie them with a piece of string (see Figure 2). Take care not to touch the soda lime; it is very caustic. Cut the loose ends of cheese cloth and string with a pair of scissors. Note: If too much soda lime is used, the bag will not fit through the neck of the flask. Prepare a second bag of soda lime similar to the first one.



*a double thickness of*
*cheesecloth containing a*
*small amount of soda lime......tied with string......loose ends cut.*

FIGURE 2: *Preparing the bag of soda lime.*

4. Record the room temperature.

5. Place a mouse in each flask. Add a little mouse food to each flask to keep the mice occupied.

6. Place a bag of soda lime in the flask that will be attached to the pipet side of the manometer. Stopper the flasks securely, but leave the vent tubes open. Note: It is very important that there by no leaks in the system while measurements are being taken. Check all connections thoroughly.

7. Wait at least two minutes for the temperature to stabilize inside the flasks. Then close the vent tubes tightly. Record the level of the manometer fluid and the time.

8. When six minutes have passed, record the new level of the manometer fluid and the time. Open both vent tubes. Calculate and record the difference between the two manometer readings.

9. Open the flasks and remove the mice. They may be returned to their cage temporarily, but make sure that you can distinguish the experimental mouse from the control mouse later.

10. Remove the bag of soda lime from the experimental flask. Allow both flasks to "air out" for five minutes.

11. Add water and ice to the water bath until the temperature is between 5 and 10 °C. Record the temperature of the water bath.

12. Return the mice to their proper flasks. Clamp the neck of each flask to a ring stand.

13. Immerse the flasks in the water bath by lowering the clamps on the ring stands. The ring stands will prevent the flasks from floating. The bath should contain enough water to submerge half to three-fourths of the flasks.

14. Repeat Steps 6 through 8 using the second bag of soda lime. Record all measurements. Add ice to the water bath if the temperature rises above 10 °C.

DISCUSSION QUESTIONS:

1. How did the change in the manometer level compare at the two different temperatures?

2. Discuss the results of the two trials in terms of the experimental mouse's elimination of carbon dioxide and consumption of oxygen.

3. In what way is the mouse's response to lowered temperature beneficial to it?

4. Describe how one could go about determining the difference in the gas pressure on the two sides of the manometer at the end of a trial.

SAMPLE DATA SHEET:

| | Room Temperature _____°C | | Water bath Temperature _____°C | |
|---|---|---|---|---|
| | time | ml | time | ml |
| Initial Reading | | | | |
| Final Reading | | | | |
| Difference | | | | |

## LABORATORY ACTIVITY 18:

## MEASURING FORCED EXPIRATORY VOLUME (FEV)

### INTRODUCTION:

An adult at rest inhales and exhales approximately half a liter of air with each breath. The breathing rate is from 12 to 18 breaths per minute. This means that the rate of ventilation is from 6 to 9 liters of air per minute. However, the maximum capacity for breathing during short intervals of extreme demand is some thirty times the ventilation rate at rest. Under such conditions, the amount of air inhaled and exhaled by a healthy person can range between four and seven liters with each breath.

This activity is divided into two parts. In Part I, the vital capacity is measured. This is also called the forced expiratory volume, or FEV, and refers to the maximum amount of air that can be expired from the lungs. In Part II, measurements are made on various parts of the body to see whether these body dimensions are related to lung volume and, if so, how.

The instrument used for measuring lung volumes is called a spirometer. Spirometers are commonly used in health testing programs. As shown on the next page, the apparatus consists of two cylindrical containers. One of these is inverted and placed inside of the other. When air is expired into the mouthpiece, the inner container rises an amount that is proportional to the volume of expired air. The volume of air expired may be read directly off of the scale mounted on the side of the spirometer.

Your data sheet for this activity will be collected by your science instructor and transferred to your mathematics instructor. The data sheet will be returned to you for further analysis in the mathematics class. For this reason, it is important that all data be clearly identified and that your name be on your data sheet.

## PART I: MEASURING VITAL CAPACITY

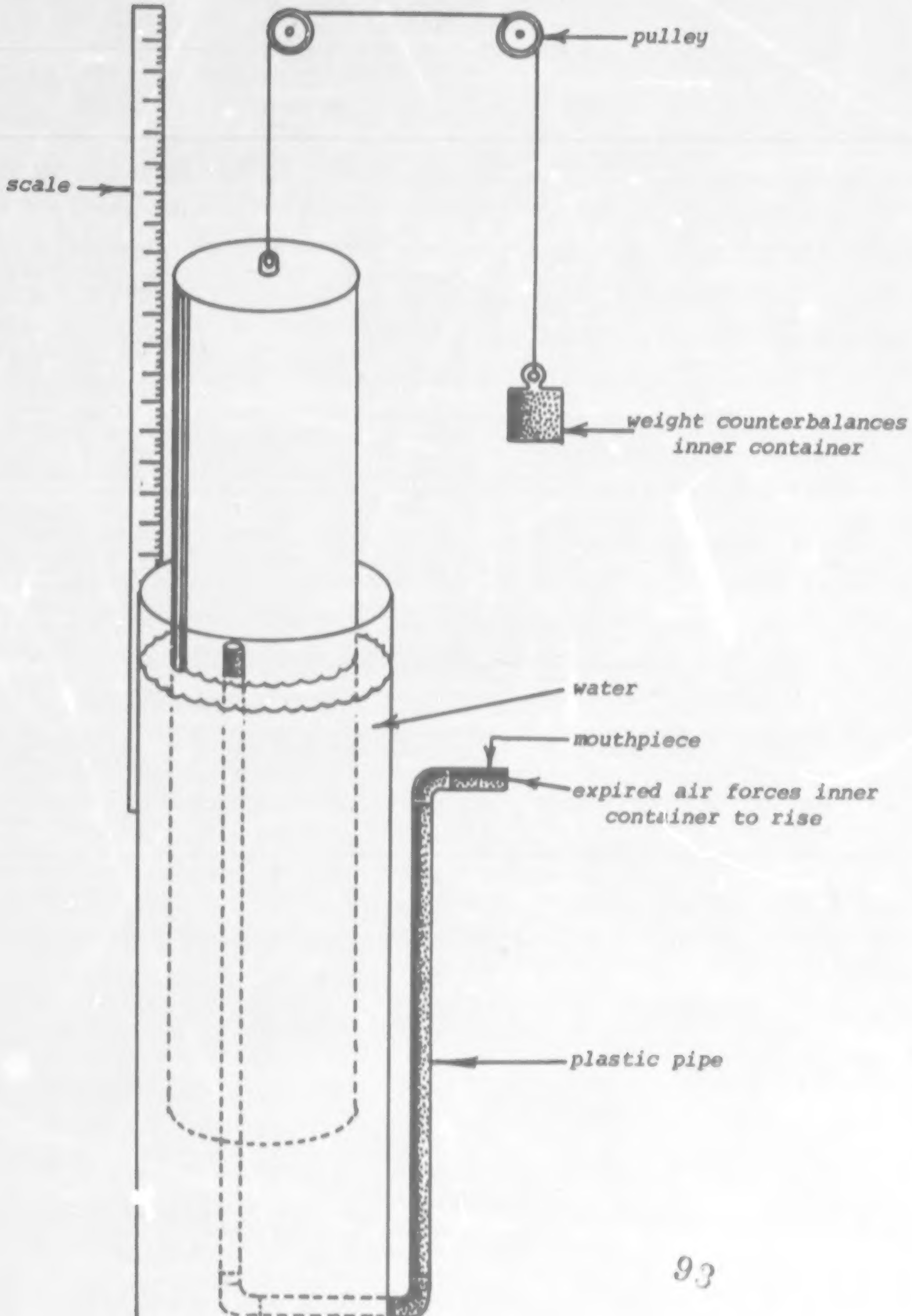
### MATERIALS:

spirometer
plastic (saran) wrap, or other suitable mouthpiece covering

### PROCEDURE:

1. Study the spirometer to make sure that you understand how it works.

# SPIROMETER


pulley
scale
weight counterbalances inner container
water
mouthpiece
expired air forces inner container to rise
plastic pipe

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Before making any measurements, always adjust the inner container to a level that corresponds to zero.

2. To maintain sanitary conditions, wrap a piece of plastic wrap around the mouthpiece of the spirometer.

3. Take as deep a breath as possible. Place your mouth over the mouthpiece and exhale into the spirometer until your lungs are emptied as much as possible.

4. Read the volume of air exhaled off of the scale mounted on the spirometer. Record the result in ml.

5. Take and record two more measurements of vital capacity in the same manner.

## PART II: MAKING BODY MEASUREMENTS

MATERIALS:

- meter stick
- tape measure (or string)
- bathroom scale

PROCEDURE:

1. Have another student measure your height to the nearest 0.01 meter. Record the result.

2. Weigh yourself to the nearest pound using the bathroom scale. Record the result. Convert this value to kilograms as follows.

$$\_\_\_\_ \text{ lb} \times \frac{1 \text{ kg}}{2.2 \text{ lb}} = \_\_\_\_ \text{ kg}$$

Show your calculations on your data sheet.

3. Have someone measure the circumference of your chest. This measurement may be taken with a tape measure, if one is available. Otherwise, a piece of string may be used. After taking the measurement, the string can be placed next to a meter stick in order to obtain the proper value. Record the result to the nearest 0.01 meter.

4. Have someone measure your arm length. The easiest way to do this is to insert the "zero" end of a meter stick inside the sleeve with the arm extended horizontally and the hand palm down. Record, to the nearest 0.01 meter, the distance to the tip of the middle finger.

5. Record your age in months.

6. Record your sex, and answer the question shown on the sample data sheet pertaining to the degree of physical activity.

7. Your instructor will indicate the procedures to be used for comparing the recorded data.

## DISCUSSION QUESTIONS:

1. Of the various kinds of data collected, which appear to be most closely related to FEV? Which appear to be unrelated to FEV?

2. Do some measurements appear to be related to FEV for one sex and not the other? If so, which ones?

3. Can you think of reasons why some of the measurements taken in Part II should be more closely related to FEV than others?

4. If the scale attached to the spirometer got lost or destroyed, what single piece of information would you need to know in order to make a new one?

## SAMPLE DATA SHEET:

Forced expiratory volume (FEV):

Trial #1 __________ ml

Trial #2 __________ ml

Trial #3 __________ ml

Height __________ m

Weight __________ lb

Weight (calculated) __________ kg

Chest circumference __________ m

Arm length __________ m

Age __________ months

Sex __________

Do you engage in vigorous physical activity 3 or more times a week? YES NO (circle one)

Is your name on your data sheet?

## LABORATORY ACTIVITY 20:

## STATIC ELECTRICITY

### GENERAL INTRODUCTION:

Have you ever walked across a carpeted room and then received a shock when you touched a doorknob? Or have you ever slid across the seat of a car and felt a shock when you grabbed the door handle? Maybe you have heard crackly sounds while combing or brushing your hair.

In these experiences, static electricity is responsible for the shocks and crackles. All ordinary matter contains both positive charges and negative charges, and it is the interaction of these two types of electric charge that accounts for the phenomena of static electricity.

Under ordinary conditions, an object contains equal numbers of positive and negative charges. The effect of one type of charge cancels out the effect of the other, and the object is said to be electrically uncharged, or neutral.

Under some conditions, however, an object may have more of one type of charge than the other. Such an object is said to be positively charged or negatively charged, depending upon whether it has an excess of positive charges or an excess of negative charges.

In this activity you will observe the effects of electric charge under a number of different circumstances. The procedure contains three parts. In Part I you will rub various objects together and investigate the resulting charges on them. Then, in Part II, you will perform a series of activities related to static electricity at various stations. Finally, in Part III, you will search for an explanation of some of the observations made at the stations in Part II.

## PART I: POSITIVE AND NEGATIVE OBJECTS

### INTRODUCTION:

When two objects are rubbed together, sometimes an excess of negative charge ends up on one object and an excess of positive charge on the other. This happens because negatively charged particles, called electrons, are rubbed off the surface of one object and onto the other.

In Part I, you will rub objects of different kinds with cloth to produce excess charge. Then you will test the charged objects to see whether their charge is positive or negative.

## MATERIALS:

2 plastic rulers (other small plastic objects may be substituted)
2 pieces of cloth
rod, 20 to 30 cm in length
glass pipet, 10-ml
ring stand
ring-stand clamp
tape
string

## PROCEDURE:

1. Suspend a plastic ruler from the rod as shown in Figure 1. Use tape to attach the string to the rod and ruler. The ruler should be balanced so that it hangs freely and can spin around without touching anything.

2. Pick up the second plastic ruler and bring one end near, but not touching, one end of the suspended ruler. Record on your data sheet whether the suspended ruler moves in response to the other ruler. Check to be sure that any movement observed is not caused by a draft in the room.

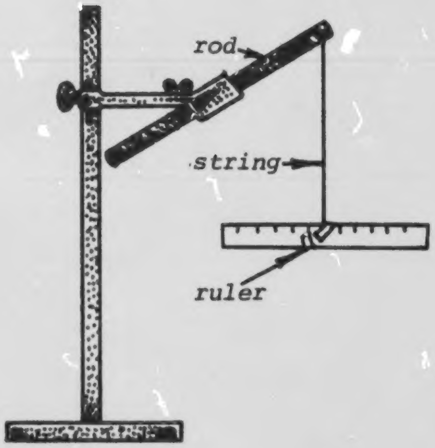


FIGURE 1:
*Suspending a plastic ruler.*

3. It is best to have a second person help you with this step. Have one person rub the suspended ruler from one end to the other with a piece of cloth, while the other does the same thing with the second ruler. Bring the rubbed end of the free ruler near one end of the suspended ruler. Record your observations. (After the suspended ruler has been rubbed, allow it to come to rest without touching it again. Otherwise, the results may be affected.)

4. Neutralize the suspended ruler by touching it with your hands. Speculate on the reason that touching the ruler eliminates the excess charge. (1)

5. Rub the suspended ruler with a cloth and take the cloth away. Then return the cloth to a position near one end of the ruler. Record your observations.

6. Touch the ruler to neutralize it. Then bring one end of a glass pipet near one end of the ruler. Record your observations.

7. Repeat Step 3, except use the pipet in place of the second ruler. Record your observations.

8. Studies have shown that when plastic is rubbed by cloth, electrons from the cloth are left on the plastic. Given this information, state the charge on the plastic (2) and on the cloth (3) after rubbing.

9. Two identical objects that are rubbed by the same material may be assumed to become charged in the same way. What is the result of bringing two objects which have the same kind of charge near one another? (4)

10. What is the result of bringing a positively charged object near a negatively charged object? (5)

11. State whether the charge on the following objects is positive or negative.

   a. Plastic ruler, after rubbing with cloth. (6)

   b. Cloth, after rubbing plastic ruler. (7)

   c. Glass pipet, after rubbing with cloth. (8)

   d. Cloth, after rubbing glass pipet. (9)

## PART II: ELECTROSTATIC PHENOMENA

### INTRODUCTION:

In this part of the activity you will have an opportunity to study several different phenomena related to static electricity. The parts will be done at different stations so that there will not be a lot of waiting. Rotate from station to station until you have completed all the activities.

### MATERIALS:

| Station | Topic | Materials |
|---|---|---|
| 1 | The Electroscope | electroscope<br>plastic ruler<br>glass pipet<br>cloth |
| 2 | How Are You Charged? | ring stand set-up from Part I<br>section of carpeting |
| 3 | Flowing Water | plastic ruler<br>glass pipet<br>cloth<br>water faucet |
| 4 | Paper the Wall | sheet of paper<br>cloth |

| Station | Topic | Materials |
|---|---|---|
| 5 | Popping Paper Confetti | acetate sheet, about 15 x 20 cm<br>5 or 6 pieces of paper confetti<br>cloth |
| 6 | Dancing Syringes | scissors<br>tissue paper<br>pie pan or other shallow container<br>plastic wrap or acetate to cover pan<br>plastic ruler<br>glass pipet<br>cloth |

PROCEDURE:

Station 1: The Electroscope

The metal-leaf electroscope is an instrument often used in the study of static electricity. As shown in Figure 2, it consists of a glass container with a stopper. Through the stopper runs a metal rod with two thin metal leaves at the bottom. When a charged object is brought near the top of the metal rod, the leaves respond.

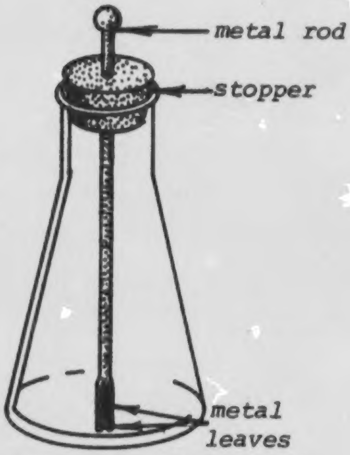


FIGURE 2:
*The electroscope.*

1. Charge an object negatively. Bring it near, but not touching, the top of the metal rod. Observe and record the response of the metal leaves.

2. Repeat Step 1 with a positively charged object.

Station 2: How Are You Charged?

The experiences described in the general introduction indicate that your body can become electrically charged. But is your charge positive or negative?

1. While standing on the carpet section, bring your finger near one end of the suspended plastic ruler. Record your observations.

2. Have someone charge the ruler negatively and repeat Step 1.

3. Rub your feet back and forth on the carpet many times, while someone is charging up the ruler again. Then repeat Step 1.

4. Replace the plastic ruler with the glass pipet. Then repeat Step 3.

Station 3: Flowing Water

Water flowing out of a faucet can be affected by a charged object. At this station, you will find out how.

1. Turn on a water faucet so that it provides a steady stream of water. Adjust the flow so that the stream is as small as possible. (The water can separate into drops lower down, as long as you have a stream 7 to 10 cm long near the faucet.)

2. Bring an uncharged object near the stream of water, and record your observations.

3. Repeat Step 2, first with a negatively charged object and then a positively charged object.

Station 4: Paper the Wall

Static electricity can cause paper to stick to a wall without the aid of glue, paste or pins.

1. Place a sheet of paper flat against a vertical surface such as a wall, door or cabinet. Let go of the paper and record your observations.

2. Repeat Step 1, but this time rub the paper several times with a cloth before letting go.

3. Lift one corner of the paper and let go of it. Record your observations.

Station 5: Popping Paper Confetti

This activity concerns the effect of static electricity on paper confetti. The explanation of what you observe will require careful observation and reasoning.

1. Place the acetate sheet on the table and hold it as flat as possible.

2. Have another person sprinkle several pieces of paper confetti on the acetate.

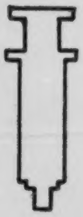
3. Lift the acetate sheet off the table and observe what happens to the bits of paper. Record your observations.

4. Repeat Steps 1 through 3, except this time rub the acetate several times with a cloth before the confetti is sprinkled on it.

Station 6: Dancing Syringes

Static electricity can be used to create some rather unusual effects, as you will see at this station.

1. Cut four paper syringes, about the size of the drawing, out of tissue paper or other light-weight paper.



2. Place the syringes in the container, and cover the container with either plastic wrap or acetate. (If you use plastic wrap, you will need enough to hold the edges together underneath the container.)

3. Rub the plastic wrap or acetate with a cloth. Observe and record the behavior of the paper syringes.

4. Bring either a positively or negatively charged object near the surface of the plastic wrap (or acetate). Observe and record the behavior of the paper syringes.

## PART III: ARE YOU POSITIVE OR NEGATIVE?

Look at the observations on your data sheet for Station 2. When you brought your finger near the negatively charged plastic ruler, you probably observed an attraction. What does this suggest about the charge on you? (10)

When you brought your finger near the positively charged glass pipet, you again probably observed an attraction. What does this observation suggest about your charge? (11)

Do you notice a contradiction between your answers to Questions 10 and 11? Could you really have been charged differently each time?

To answer these questions, consider your observations at Station 4. When you rubbed the paper, it probably stuck to the wall. Why? We know that positively and negatively charged objects attract each other, so this could explain it. But the wall had no charge at all to start with--it wasn't rubbed. So why did the paper stick to it?

When you rubbed the paper it became negative. The reason the paper stuck onto the wall is that the negative charge on the paper caused a positive charge in the wall, and then the paper was attracted to the wall. Here's how it works.

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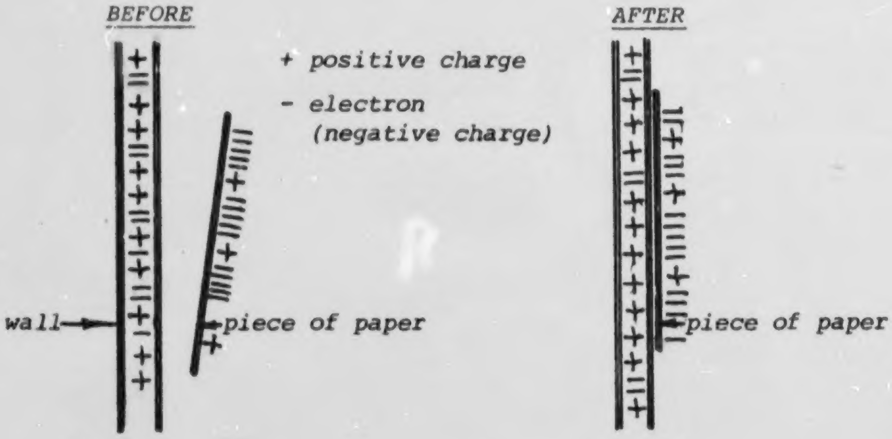
The wall had no charge--it had a normal number of electrons. Negative charge, such as the paper had, means there are excess electrons. All these excess electrons on the paper repelled the electrons in the part of the wall behind the paper, which left that part of the wall deficient in electrons. (The repelled electrons moved away to the rest of the wall.)

A deficiency of electrons at the part of the wall behind the paper meant a positive charge there. That part of the wall was thus positive. This is why the negative rubbed paper was attracted to it. Figure 3 shows this process, called induction. The negative paper induces a positive charge on the wall.



*The wall has no charge at first. It has a normal number of electrons. The piece of paper has a negative charge. It has an excess of electrons.*

*The excess electrons on the paper repel the elctrons that are on the wall, leaving the part of the wall behind the paper deficient in electrons. This results in a positive charge on that part of the wall. Therefore the negative paper is now attracted to the positive part of the wall.*

FIGURE 1: *Induction, or why a charged object can be attracted to one with no charge.*

Now we can return to the apparent contradiction at Station 4, where you probably found that your finger attracted both a positive and a negative object. The answer lies with induction. To understand how, you may need to know one more fact: when you rubbed your feet on the carpet you became negative. In fact, you attained a very powerful negative charge, far greater than the charges on the small ruler and pipet.

You had a powerful negative charge. This means you had a large excess of (12) (electrons, positive charges). Excess electrons on your finger (13) (attracted, repelled) the electrons at one end of the plastic ruler. This left a (14) (positive, negative) charge at that end of the ruler. Thus that end of the ruler was attracted to your negative finger, even though the rest of the ruler had a slight negative charge.

Induction also contributed to attracting the positively charged pipet. It helped to make the end even more positive.

Induction also explains observations you may have made at Station 3. You probably found that both positive and negative objects attracted the stream of water. Now, answer questions 15 through 17.

When a negative object was brought near the stream, by induction the object caused a (15) (positive, negative) charge in the water. This (16) (attracted, repelled) the water.

When a positive object was brought near the stream, by induction it caused a (17) (positive, negative) charge in the water. This attracted the water.

DISCUSSION QUESTIONS:

1. Explain the behavior of the electroscope when charged objects are brought near it. (Hint: when do two objects repel one another? Also, consider induction.)

2. Explain the electroscope's behavior when the charged object is removed.

3. Use the information you have gained to explain, as well as you can, your observations at Station 5 (Popping Paper Confetti) and Station 6 (Dancing Syringes).

## LABORATORY ACTIVITY 22:

## ELECTROPLATING

INTRODUCTION:

In today's activity you will perform a process known as electroplating. In electroplating, electric current is used to cause a coating of metal to be deposited on an object. Many of the metal objects you encounter every day are electroplated. Spoons, watchcases, car bumpers and many other objects are coated with various metals (silver, gold, chromium) to prevent rusting and/or to improve their appearance.

How does electroplating work? Hopefully, when you finish this activity you will be able to answer this question. However, you may need a few clues. For example, electroplating makes use of two concepts you have studied recently. (1) Ionic compounds dissociate in solution into their independent components or ions and these ions carry a charge, either positive or negative. (2) Unlike charges attract each other and like charges repel each other.

In electroplating, an electric current is caused to flow between two points, or terminals. This is accomplished by using a voltage source (the BIP). As you may recall from Section 20, voltage may be thought of as the force that causes the separation of negative and positive charges. By hooking two terminals up to a voltage source one terminal can be made to accumulate negative charge (this is the cathode) and the other terminal positive charge (this is the anode). You will be hearing more about the anode and the cathode in the future. If a conductor is placed in between the cathode and the anode, current will flow. Negative charges will flow from the cathode to the anode and positive charges will flow from the anode to the cathode (see Figure 1).

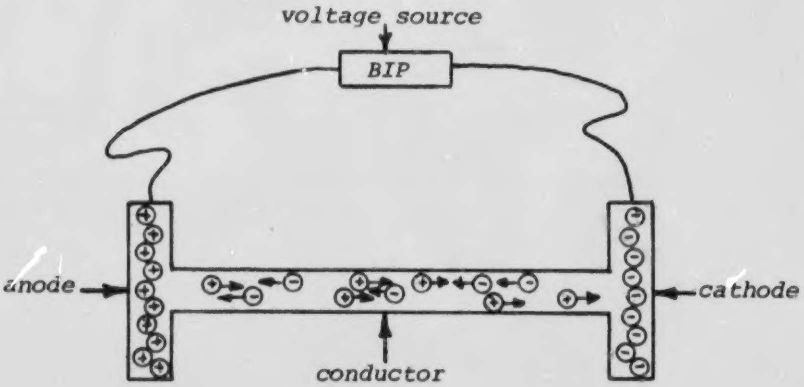


FIGURE 1: *Flow of electric current between anode and cathode.*

What type of metal coating is deposited on an object in the electroplating process depends on the choice of the conductor in Figure 1. In this activity the conductor is a solution of copper sulfate ($CuSO_4$) dissolved in water. Copper sulfate is an ionic compound that dissociates into copper ions ($Cu^{+2}$) and sulfate ions ($SO_4^{-2}$).

As you set up the equipment and perform the activity try to visualize what is happening in terms of the movement of electric charges. This will help you answer the questions at the end of the activity.

MATERIALS:

100 ml sulfuric acid solution
20 g $CuSO_4$ crystals
graduated cylinder, 100-ml
beaker, 250-ml
BIP and programming wires
2 wires with alligator clip on one end
10kΩ resistor
2 carbon rods
glass stirring rod
glass-marking pencil
metric ruler
balance
plastic container
spatula or scoopula
steel wool
tape
soap

PROCEDURE:

1. As stated in the introduction, you will be using the BIP as a voltage supply to create the electric current necessary for the electroplating. Connect the 5-volt supply on the BIP as indicated by the instructor. When you are finished you should have a set-up like the one shown in Figure 2. (Note: DO NOT plug in the BIP until indicated in the procedure.)

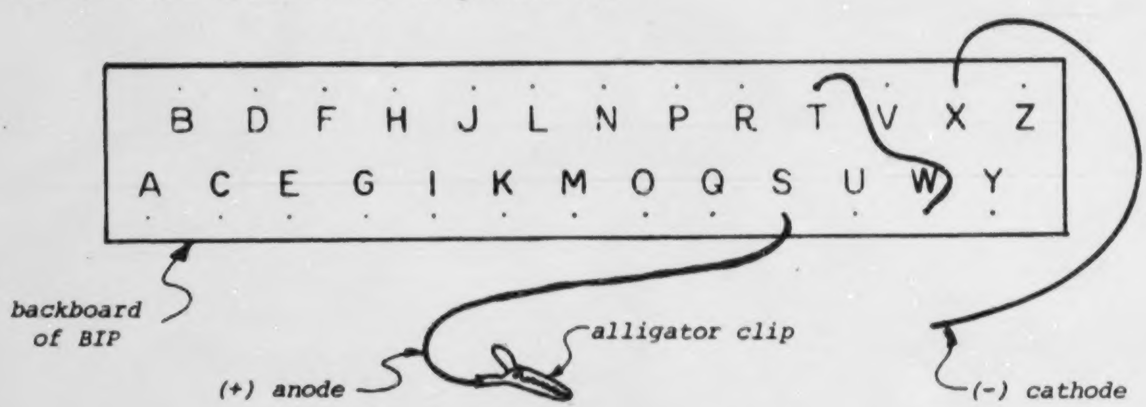


FIGURE 2: *BIP programming for 5-volt supply.*

2. Hooking up the 5-volt supply on the BIP means that there will be a 5-volt drop (remember the waterfall in Section 20) between the anode (+) and the cathode (-). For the electroplating you are going to do it is desirable that the voltage drop be closer to two volts. In order to accomplish this a resistor (something that resists the passage of electric current) is placed in the circuit. The resistor lowers the amount of current that passes through the circuit. As was discussed in Section 20, a lower current means the voltage will also be lower. In this particular case, it is necessary to add a resistor of 10,000 ohms to obtain the voltage drop of two volts that is wanted. (Ohms are merely the units in which resistance is measured.) Clip in the 10,000 ohm resistor as shown in Figure 3.

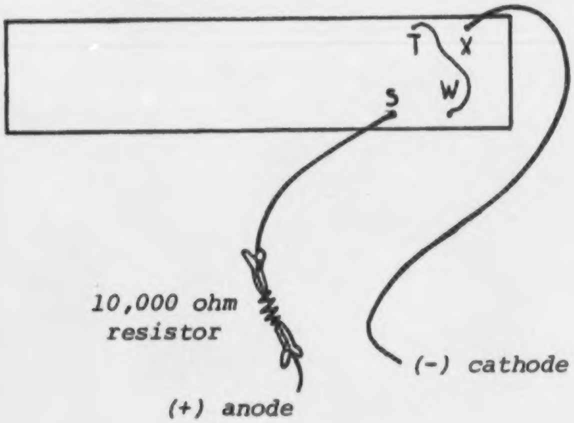


FIGURE 3: *Circuit with resistance added.*

3. The next step is to attach one carbon rod to the anode and another to the cathode as shown in Figure 4 on the following page. Before doing this clean each rod carefully with steel wool and soap. Rinse well. (It is important that the surface be free of foreign particles and grease or the plating will not be uniform.) Also, to assure a good connection, the wire should be wound around the carbon rod three or four times. You may need to peel away part of the insulation to expose sufficient wire.

4. You now have the voltage source, resistor and terminals connected. It remains for you to prepare the conductor, i.e., the copper sulfate ($CuSO_4$) solution, and immerse the terminals in it. To prepare this solution weigh out 20 ± 0.1 grams of copper sulfate crystals and place them in the container you are going to use for the electroplating. Measure out 900 ± 10 ml of hot tap

water, add this to the container and stir with a glass rod until the $CuSO_4$ has dissolved. (The $CuSO_4$ crystals dissolve more rapidly in hot water.)

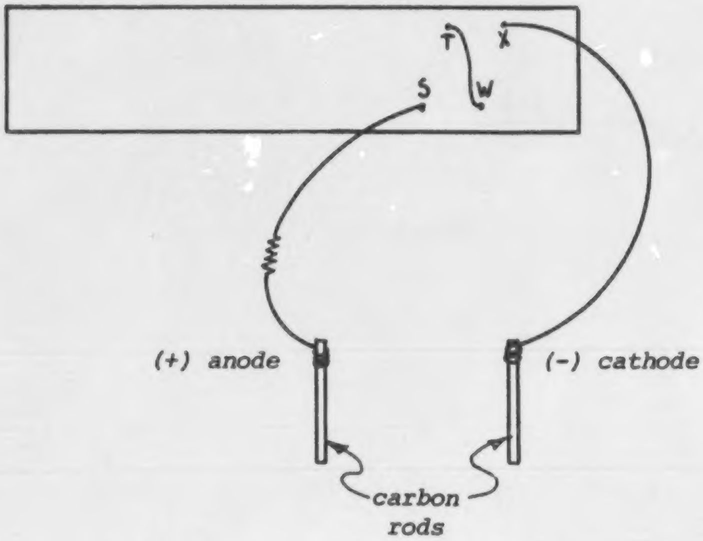


FIGURE 4: *Attachment of the carbon rods.*

5. Add 100 ± 1 ml of the sulfuric acid solution. (The acid is added to improve the quality of the plating.) CAUTION: Handle the acid with care. It can burn the skin and ruin clothing. If you should get acid on your hands, don't panic. Wash it off with plenty of water, and no harm will be done.

6. Next, mark two points on the middle portion of the glass stirring rod about 10 ± 1 cm apart. Fix the anode and cathode to these points by coiling the insulated wire around the rod and taping it in place.

7. When you have attached the two terminals lay the glass rod over the top of your container so that the carbon rods hang down into the solution. Most of the rods should be immersed, but not the area where they are connected to the wires leading to the BIP. Your completed set-up should look like that in Figure 5, on the next page.

8. Plug in the BIP and let the electroplating proceed for ten minutes.

9. Unplug the BIP and remove the rod and terminals from the solution. Do not remove the carbon rods from the glass rod. Rather set the whole apparatus down on a clean paper towel.

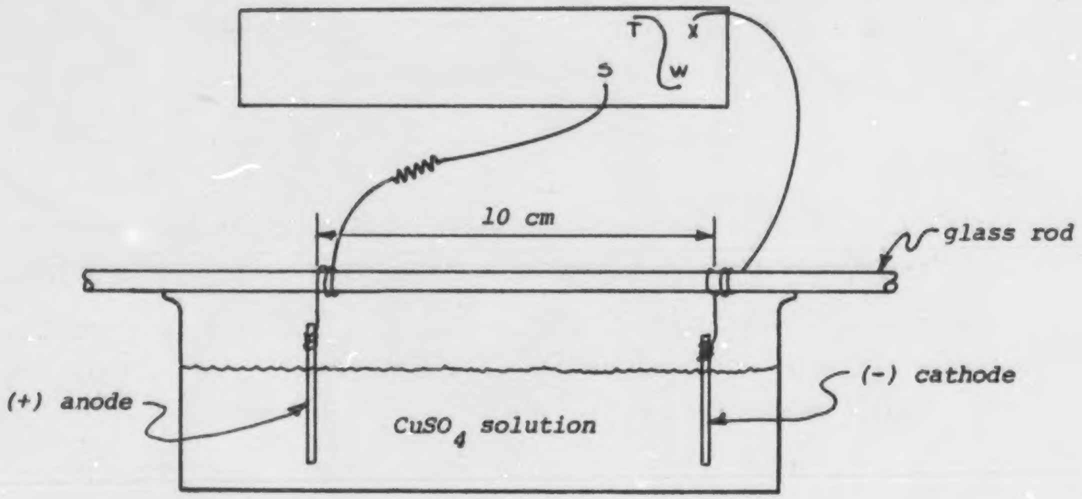


FIGURE 5: *Completed electroplating set-up.*

DISCUSSION QUESTIONS:

1. When the current travels through metal the carrier of electric charge is the electron. When electric current travels through a liquid what are the carriers? What were the carriers in the activity you just performed?

2. Why does a copper coating form on the rod at the cathode (-) and not at the anode (+)? Can you explain this in terms of the two concepts ennumerated in the second paragraph of the introduction?

3. How does the copper dissolved in the solution differ from the copper deposited at the cathode? Explain the difference in terms of atomic structure.

4. In this activity you used electric current to cause copper to be deposited on a carbon rod at the cathode. Can you suggest a way to use the same principles to cause the copper coating to be removed from the carbon rod? Explain your reasoning in terms of electric charges.

Test your idea. Use the electroplating equipment you have to remove the copper plating from the carbon rod. Draw a diagram of the set-up you used and describe your results.

## LABORATORY ACTIVITY 23:

## MOLECULAR MODELS

### INTRODUCTION:

In Section 23 we introduced three types of formulas used to represent molecules: electron-dot formulas, structural formulas and molecular formulas.

Molecules may also be represented using balls as models for atoms and sticks or springs for bonds. In this activity you will be asked to construct models of a number of different molecules.

Your instructor will indicate which types of balls should be used for which elements. Note that the ball used as hydrogen has one hole; a hydrogen atom has one electron and needs to share one more to attain the electron configuration of the noble gas helium. A ball representing carbon has four holes; a carbon atom needs four electrons to complete its outer shell. A nitrogen ball has three holes, oxygen balls and sulfur balls have two holes, while balls used for chlorine have one hole.

In constructing the models, there are three rules to be followed.

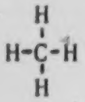
1. Each ball must have all its holes filled. For example, each carbon atom must have four sticks or springs attached to it.

2. Each stick or spring you use must have a ball attached on both ends.

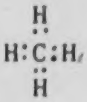
3. No two adjacent balls may have more than one stick or spring connecting them. (This rule will be broken in the future.)

### EXAMPLE:

$CH_4$--methane



$$\begin{array}{c}H\\ |\\ H-C-H\\ |\\ H\end{array}$$



$$\begin{array}{c}H\\ \cdot\cdot\\ H:C:H\\ \cdot\cdot\\ H\end{array}$$

### MATERIALS:

H--10

C-- 4

Cl--3

O--2

N--1

S--1

13 sticks or springs

### PROCEDURE:

Construct models of the following molecules. When you decide that a model has been constructed correctly, write the structural formula and electron-dot formula of the molecule on your data sheet.

1. HCL--hydrogen chloride
2. $NH_3$--ammonia
3. $H_2O_2$--hydrogen peroxide
4. $H_2S$--hydrogen sulfide
5. $CHCl_3$--chloroform
6. $CH_4O$--methyl alcohol
7. $C_2H_6O$--ethyl alcohol

   $C_2H_6O$--methyl ether

   (Two structures are possible. Construct both. There is no reason for you to know which is ethyl alcohol and which is methyl ether at this time.)
8. $CH_5N$--methylamine
9. $C_4H_{10}$--butane

   $C_4H_{10}$--isobutane

   (Construct both.)

## LABORATORY ACTIVITY 24:

## CONDUCTIVITY OF SOLUTIONS

### INTRODUCTION:

The ability of certain kinds of liquids to conduct electricity is of fundamental importance to the existence of human life. For example, the transmission of nerve signals between the brain and other parts of the body, the function of the muscles, and even the ability of the heart to pump blood are all dependent upon the fact that body fluids conduct electrictiy. The examples just mentioned will all be discussed in later parts of the course, but first it is important to examine the features of liquids that contribute to their ability to conduct electric current. This ability to conduct current is called conductivity.

In Laboratory Activity 22 you observed one example of the conductivity of a solution. The process of copperplating depends upon the ability of the copper sulfate solution to conduct electricity. The plating process cannot work unless the solution permits electrons to move from the cathode to the anode.

In this activity, you will have an opportunity to compare the conductivities of some pure solvents and to test the effect upon conductivity of dissolving different quantities of various solutes in water. The set-up will be almost the same as you used for copper plating. The BIP will be used to provide electric current to two carbon electrodes. In this case, however, the circuit will contain a small light bulb. When the two electrodes are placed in a test solution, the brightness of the bulb will give an indication of the conductivity of the solution (or lack of it).

The solutions to be tested are given in the materials list. The values given indicate the quantity of solute that was used to make one liter of the test solution.

### MATERIALS:

HCl (hydrochloric acid)--
4 grams/liter

HCl (hydrochloric acid)--
0.4 gram/liter

NaOH (sodium hydroxide)--
4 grams/liter

NaOH (sodium hydroxide)--
0.4 gram/liter

$CuSO_4$ (copper sulfate)--
80 grams/liter

$CuSO_4$ (copper sulfate)--
16 grams/liter

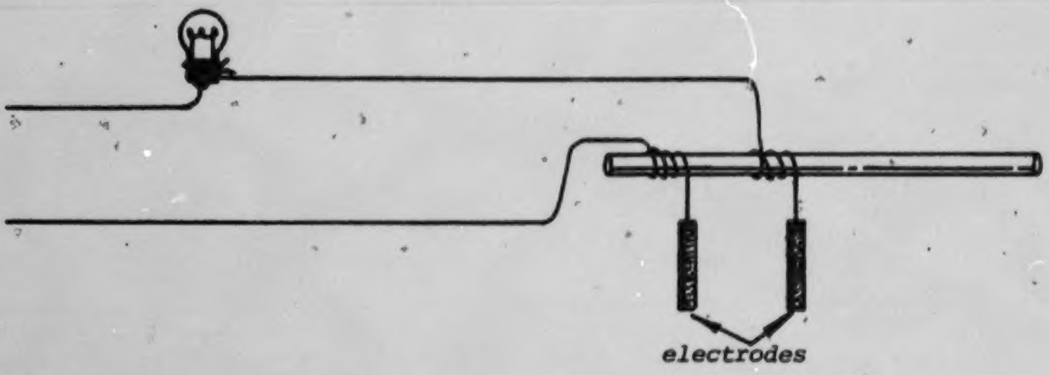
NaCl (sodium chloride)--
4 grams/liter (table salt)

$C_{12}H_{22}O_{11}$ (sucrose)--80 grams/liter
(table sugar)

denatured alcohol
beaker, 250-ml
tap water
BIP and programming wire
wire cutter-strippers
screwdriver
light bulb attached to programming wire
2 carbon electrodes attached to programming wires
stirring rod

PROCEDURE:

1. Attach the two electrodes to the stirring rod by wrapping a few coils of wire around the rod as shown in the figure. The electrodes should be near one end of the rod and close enough together to fit inside a 250-ml beaker. Use the wire stripper to bare about 4 cm of one of the electrode wires. Wrap the wire around the threads of the light bulb and twist it to obtain a secure connection.



2. Program the BIP by connecting the two wires from the light bulb and electrode to S and W. (It doesn't matter which is which.) Connect T to U. Notice that you can hold one end of the stirring rod and immerse the electrodes in a test solution without touching either the electrodes or the solution. This way you can keep yourself from becoming part of the electric circuit. In addition, four of the solutions should not be touched. Solutions containing HCl and NaOH should always be handled with care to avoid spills. If you should get either on your hands, however, don't panic. Walk calmly to the nearest faucet and wash it off with plenty of water.

3. Insert a screwdriver into the slotted control of the voltage regulator, which is below and to the left of the mA dial. Turn the control as far clockwise as possible.

4. Fill a 250-ml beaker with tap water. This beaker is to be used to rinse the electrodes after each test. It is also a good idea to replace the water after each rinse so that the test solutions do not become contaminated.

5. Obtain a test solution and plug in the BIP. Place the electrodes in the solution and observe the brightness of the light. Record your observations using a brightness scale such as the following one.

0 (no light)
+ (dim)
++ (bright)
+++ (very bright)

In order to make comparisons, it will be necessary to dip the electrodes to the same depth for each solution. (Test the effect of changing the length of the electrodes that is immersed. Does it make a difference?)

6. Continue until you have tested all solutions available. Remember to rinse the electrodes between tests.

DISCUSSION QUESTIONS:

1. Which solutions were the best conductors?

2. Which solutions were the worst conductors?

3. Does the amount of a substance dissolved appear to have an effect on conductivity? Explain.

4. Compare the conductivity of HCl, NaOH and NaCl when 4 grams are dissolved in a liter of water.

5. Which of the substances tested would you expect to have ionic bonds? Which would you expect to be covalently bonded? Why?

6. Would you expect sulfuric acid to be a good conductor? What about maple syrup and whiskey?

## LABORATORY ACTIVITY 26:

## CHEMICAL REACTIONS, A QUANTITATIVE STUDY

### INTRODUCTION:

In this activity you will be given a chemical equation similar to the ones discussed in Section 26. You will also be provided with specific proportions of the two reactants and asked to predict the amounts of the products that will be formed. You will then mix solutions containing the reactants and observe the formation of the products. After the reaction has taken place the products will be separated and weighed in an attempt to confirm your calculations. In this way you will be able to test whether the theory on which your calculations are based does indeed hold up when tested in the laboratory.

The equation that you will be using is shown below.

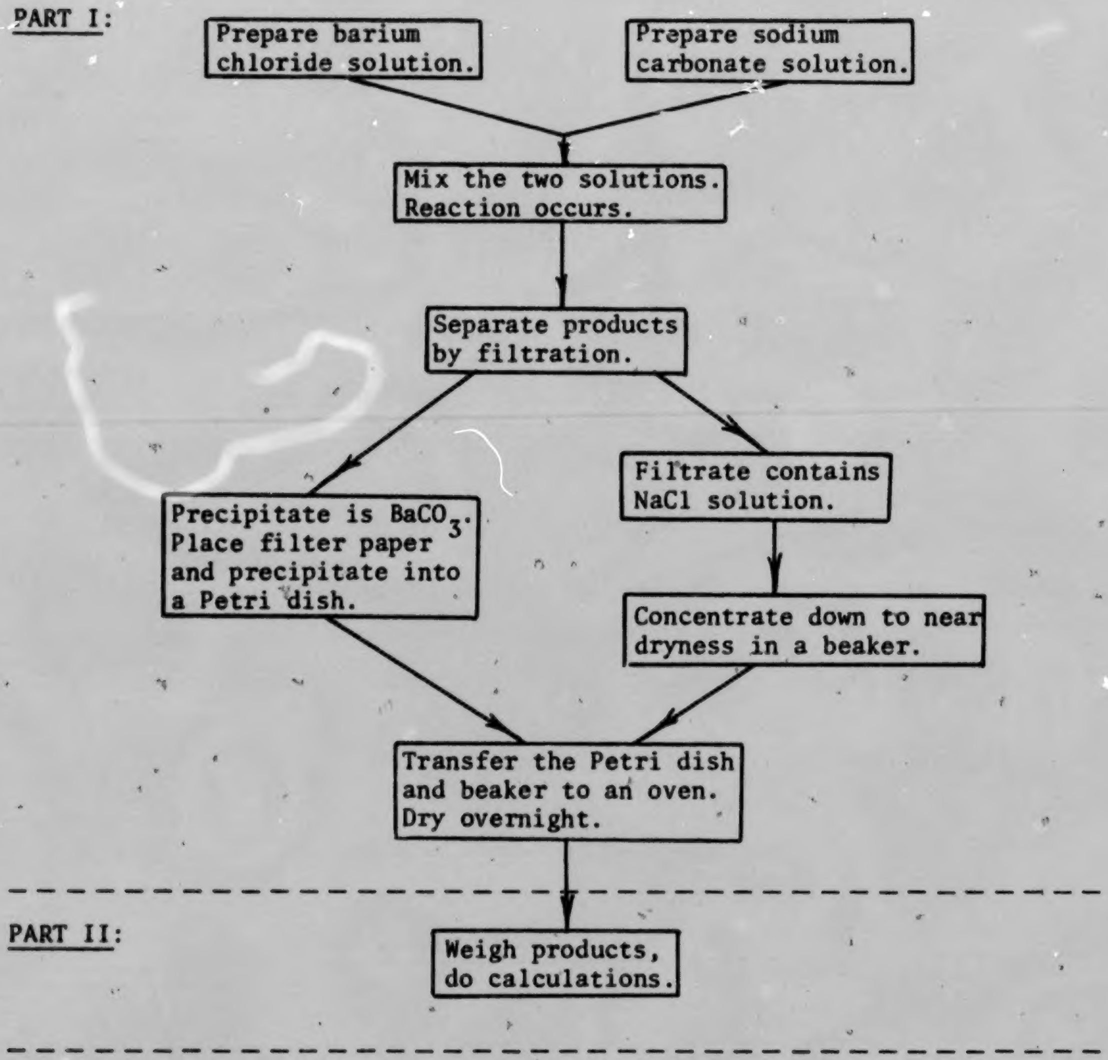
| $BaCl_2 \cdot 2H_2O$ | + | $Na_2CO_3$ | → | $BaCO_3$ | + | $NaCl$ | + | $2\ H_2O$ |
|---|---|---|---|---|---|---|---|---|
| (barium chloride) | | (sodium carbonate) | | (barium carbonate) | | (sodium chloride) | | (water) |

You may have noticed that the chemical formula for barium chloride contains a notation that you have not seen before, namely the two water molecules ($2H_2O$). This type of compound is known as a hydrate, and the two water molecules are called the water of hydration. What this means is that the barium chloride crystals you will be using have water molecules incorporated into their structure. The number preceding the formula for water (in this case, 2) indicates how many water molecules are associated with each molecule of the compound. Of course when you calculate the molecular weight of such a compound the water molecules must be taken into account. While you are performing this activity you might think about what effects, if any, the water of hydration in the barium chloride will have on your calculations and the actual results you obtain.

This activity will also introduce you to a new laboratory technique, filtration. Filtration is used to separate a solid from a liquid. In this particular activity the solid in question is the $BaCO_3$ that forms when the reaction takes place. By filtering the mixture you can separate the $BaCO_3$ crystals from the NaCl solution. The crystals will remain behind on the filter paper, and the solution will pass through the paper and be collected in a beaker. You may then crystallize the NaCl by evaporating the water from the solution. Finally, after drying the $BaCO_3$ and NaCl crystals you can weigh them to ascertain the amount of each product produced by the reaction.

The following flow chart may help you to get an overview of the entire procedure.

FLOW CHART


PART I:
Prepare barium chloride solution.
Prepare sodium carbonate solution.
Mix the two solutions. Reaction occurs.
Separate products by filtration.
Precipitate is BaCO3. Place filter paper and precipitate into a Petri dish.
Filtrate contains NaCl solution.
Concentrate down to near dryness in a beaker.
Transfer the Petri dish and beaker to an oven. Dry overnight.
PART II:
Weigh products, do calculations.

## PART I: THE REACTION OF BARIUM CHLORIDE AND SODIUM CARBONATE

**MATERIALS:**

1.06 g sodium carbonate ($NaCO_3$)
2.44 g barium chloride ($BaCl_2 \cdot 2H_2O$)
oven
3 beakers, 150-ml
graduated cylinder, 100-ml
Petri dish
filter paper
stirring rod
gas burner
ring stand, with ring
balance
wire gauze

PROCEDURE:

1. First you will need to prepare two solutions, one of each reactant. To do this, weigh out 2.44 ± 0.01 g of barium chloride hydrate and 1.06 ± 0.01 g of sodium carbonate.

2. Place the barium chloride hydrate in one beaker and the sodium carbonate in another. Add 25 ± 1 ml of tap water to each beaker. Stir each solution for a few minutes until all the crystals have dissolved. Do you think the reaction would occur if you simply mixed the crystals together without using any water?

3. Next, mix the two solutions together. Do this by pouring the contents of one beaker into the other and swirling the mixture gently a few times. Rinse out the beaker that you emptied with a very small amount of water and add this to the beaker that contains the mixture. (Why do you think this rinse is necessary?)

4. At this stage the reaction should have occurred, as evidenced by the formation of a white precipitate--this is the $BaCO_3$. The next step will be to separate the precipitate from the NaCl which is still dissolved.

5. Start by placing your initials on a clean beaker and on a filter paper. Then weigh each of them separately to the nearest 0.01 g. Record the two masses.

6. Make a filter by folding the filter paper first into a semicircle and then into a quarter circle. Open the paper into a funnel shape as shown in Figure 1.

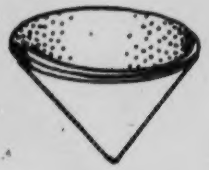


FIGURE 1: *Making a filter.*

7. Place the filter in the weighed beaker, and pour the reaction products through the filter paper as shown in Figure 2 on the next page. Again, use a small amount of water to rinse out the beaker and pour it through the filter. The filtering process will take approximately thirty minutes.

8. While you are waiting, initial and weigh a Petri dish half (either a cover or a bottom). Record the mass.

9. When the filtering is complete, take the filter paper with the precipitate and set it in the Petri dish. Place the Petri dish and the filter paper in an oven for drying.

10. Take the beaker that contains the filtrate (the liquid left over after the filtering is complete) and use a burner to boil off almost, but not all, the water.

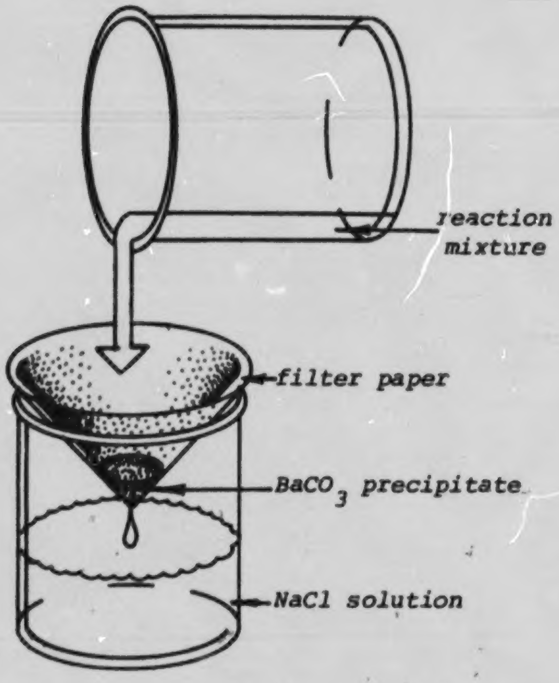


FIGURE 2: *Filtration of reaction mixture.*

11. Place the beaker containing the NaCl crystals in the oven. They will remain there overnight, until the next lab period, when you will weigh the dried products.

PART II: PRODUCT MASSES AND CALCULATIONS

MATERIALS:

balance

PROCEDURE:

1. Remove your Petri dish with the filter paper and your beaker from the oven. Allow them to cool to room temperature. Weigh the beaker and then the Petri dish containing the filter paper. Record the masses.

2. Determine the masses of the products and record them.

CALCULATIONS:

1. What are the gram molecular weights of $BaCl_2 \cdot 2H_2O$, $Na_2CO_3$, $BaCO_3$ and NaCl?

2. What is the conversion factor for calculating how many grams of $BaCO_3$ are formed per gram of $BaCl_2 \cdot 2H_2O$ reacting?

3. What is the conversion factor for calculating how many grams of NaCl are formed per gram of $Na_2CO_3$ reacting?

4. Given 2.44 g of $BaCl_2 \cdot 2H_2O$ reacting with $Na_2CO_3$, how many grams of $BaCO_3$ should be formed?

5. Given 1.06 g of $Na_2CO_3$, how many grams of NaCl should be formed?

DISCUSSION QUESTIONS:

1. Were your calculated masses of the products, $BaCO_3$ and NaCl, verified by the results you obtained in the activity? If not, give some possible reasons for the differences.

2. Do you think the water of hydration of the barium chloride had any effect on your calculations and your results? If so, explain why, and suggest a way of correcting for such effects.

3. Did the mass of the reactants equal the mass of the products? If not, can you explain why not?

SAMPLE DATA SHEET:

Mass in grams of:

Reactants $\begin{cases} BaCl_2 \cdot 2H_2O & \_\_\_\_\_\_ \\ Na_2CO_3 & \_\_\_\_\_\_ \end{cases}$

filter paper ______

beaker ______

Petri dish ______

beaker + NaCl ______

Petri dish + filter paper + $BaCO_3$ ______

Products $\begin{cases} NaCl & \_\_\_\_\_\_ \\ BaCO_3 & \_\_\_\_\_\_ \end{cases}$

## LABORATORY ACTIVITY 28:

## PREPARATION OF MOLAR SOLUTIONS

### INTRODUCTION:

In the clinical laboratory, tests using chemical solutions are frequently performed. The results of such tests can be used in diagnosing and treating various medical problems or conditions. These tests are typically done with solutions of known molar concentration, since compounds react in molar proportions. In doing the final calculations, the concept of the mole can simplify problems of great complexity.

Since the outcome of such tests can be of critical importance, it is essential to be able to prepare solutions of the needed concentrations. This activity provides a general procedure and technique for preparing molar solutions. The instructor will tell you which solutions to prepare, their concentrations and volumes. You will calculate the amount of substance required (in grams), and then use the procedure to prepare the solution.

These solutions are to be used in later activities. Since proper results can depend upon solutions that are correctly prepared, it is important to be careful with the calculations and in following the procedure.

### MATERIALS:

- balance
- beaker (size depending upon the volume of solution being prepared)
- stirring rod
- bottle (or flask with stopper)
- graduated cylinder (size depending upon volume of solution being prepared)
- labels
- scoopula
- samples of chemicals needed to make the solutions

### PROCEDURE:

1. Calculate the gram molecular weight of the solute.

2. Use dimensional algebra to calculate the mass of solute needed to prepare a solution of the specified concentration and volume.

3. Use the balance to weigh out the mass of solute needed to the nearest 0.01 g.

4. Transfer the sample to a clean beaker. (The beaker does not have to be dry. Can you explain why not?)

5. Add water to the beaker to dissolve the solute. The use of a stirring rod may help dissolve the sample. (Note: the quantity of water used should be at least 30 to 50 ml less than the desired final volume. The reason for this will become apparent in the following steps.)

6. Pour the contents of the beaker into a graduate. (Again, the graduate does not have to be dry, but must be clean.)

7. Add 10 to 20 ml of water to the beaker and swirl the contents briefly in order to mix any remaining drops that contain the solute. Rinse the end of the stirring rod in this solution if it was used in Step 5. (The point is to make sure that all of the original sample weighed in Step 3 is included in the final solution. This is one of the most important things to keep in mind when preparing solutions. In some cases, it may be desirable to repeat this rinse if it is felt that a significant amount of the sample still remains in the beaker.)

8. Add this solution to the contents of the graduate.

9. Fill the graduate to the level that corresponds to the volume of solution specified in the instructions. The solution in the graduate should now be the correct volume and concentration.

10. Pour the contents of the graduate into a storage bottle (or flask). It may be necessary to swirl the solution briefly (or to pour the solution into the graduate and then back into the bottle) in order to be certain that the contents are thoroughly mixed.

11. Label the bottle with the name of the chemical and its concentration. If the chemical is highly poisonous or corrosive, it should be indicated on the label. Make sure that the bottle is capped tightly.

12. SPECIAL INSTRUCTION #1: In some cases, a solution must be prepared by diluting an already existing solution of known concentration. For example, if a 0.001 M solution of sodium hydroxide (NaOH) is needed, it may be prepared by mixing 1.0 ml of 0.1 M NaOH with 99.0 ml of water. This means that the 0.1 M solution is diluted by a factor of 100 and that the final concentration is 0.001 M. (The 1.0 ml of 0.1 M NaOH would best be measured using a 1-ml pipet.)

13. SPECIAL INSTRUCTION #2: In some cases, solutions that have very small concentrations must be prepared by first making a concentrated solution and then

diluting it as explained in Step 12. For example, to prepare the 0.001 M NaOH solution described in Step 12, 0.004 g of NaOH would have to be dissolved in 100 ml total volume. This small quantity of NaOH cannot be weighed accurately with a student balance. Therefore, it is simpler and more accurate to first make a 0.1 M solution (0.4 g NaOH in 100 ml), and then to dilute by a factor of 100.

## DISCUSSION QUESTIONS:

1. How would the procedure given here have to be changed, if you needed to prepare a solution having a volume of only 50 ml?

2. Step 11 of the procedure indicates that the storage bottle should be tightly capped. Give at least two reasons for this.

## LABORATORY ACTIVITY 29:

## MEASURING THE CONCENTRATION OF ACIDS BY TITRATION

### INTRODUCTION:

In the clinical laboratory, it is often very important to be able to determine the concentration of a dissolved substance. The result may be a matter of life and death. By using the proper procedure, it is possible to measure the concentration of virtually any solution. Examples of concentration determinations to be performed in this course include the amount of Vitamin C in orange juice, the amount of salt and protein in various foods, and the amount of urea in urine.

One way to measure concentration is to combine a sample of the substance with a testing reagent of known concentration. The amount of test reagent that is required to react completely with the sample can be used to calculate the unknown concentration. One procedure of this kind is called titration. It is a volumetric test since it measures the volume of liquid required to react with a given sample. (It is also possible to measure concentration using a colorimetric test, which will be described in Laboratory Activity 31.)

The volume of test reagent required is typically measured with a buret. Along with the pipet and the graduated cylinder, the buret is one of the most common pieces of laboratory apparatus used for measuring volumes of solutions. A buret can typically release up to 50 ml of liquid with a precision of ± 0.05 ml.

The purpose of this activity is to measure the concentrations of three acids by titrating them with a basic solution of known concentration. When dissolved, an acid (represented as HA) releases hydrogen ions into solution, while a base (BOH) releases hydroxide ions.

$$HA \rightarrow H^+ + A^-$$

$$BOH \rightarrow OH^- + B^+$$

When mixed together, acids and bases react to form water (HOH) and a salt (BA).

$$BOH + HA \rightarrow HOH + BA$$

A given number of moles of hydrogen ions react with an equivalent number of moles of hydroxide ions. Therefore, the concentration of an acid solution may be determined by measuring the amount of base that is needed to react completely with the acid.

As an example, suppose that 20 ml of 0.1 M sodium hydroxide ($NaOH$) are required to react completely with 10 ml of a hydrochloric acid ($HCl$) solution. The unknown concentration of acid would be 0.2 M, that is, twice as concentrated as the base. The following formula is commonly used for doing such calculations.

$$C_A V_A = C_B V_B \quad \text{or} \quad C_A = \frac{C_B V_B}{V_A}$$

$C_A$ and $C_B$ stand for the concentrations of acid and base, respectively, expressed in moles per liter (M). $V_A$ and $V_B$ stand for the volumes of acid and base used in the titration. A sample calculation of the preceding example is given.

$$\text{HCl concentration} = C_A = \frac{(0.1\ \text{mole/liter})(20\ \cancel{\text{ml}})}{(10\ \cancel{\text{ml}})}$$

$$= 0.2\ \text{mole/liter} = 0.2\ \text{M}$$

In order to determine when the acid-base reaction is complete, a special chemical called an indicator is used. The indicator is a dye whose color depends upon the concentration of hydrogen ions. In this activity, a few drops of phenol red solution are added to each sample. The phenol red is colored yellow in acidic solution, but turns red as soon as an excess amount of bases is added. The point at which the indicator changes color is called the endpoint of the titration.

The procedure is divided into three parts. Part I provides some general instructions on the care and use of the buret. In Part II, the concentrations of hydrochloric acid and vinegar samples are calculated following titration with 0.1 M sodium hydroxide.

In Part III, the molecular weight of an unknown solid acid is determined. This is done by first dissolving 1.0 g of the unknown acid in 100 ml of water. A 10 ml sample of the resulting solution is titrated with 0.1 M sodium hydroxide. The molarity of the acid solution is calculated, which can then be used to determine the molecular weight of the unknown. For example, if the molarity of the solution of the unknown solid acid is calculated to be 0.25 M, the molecular weight of the acid can be determined as follows.

$$\text{molecular weight of unknown} = \frac{\text{mass concentration}}{\text{molar concentration}}$$

$$= \frac{(1.0\ \text{g})/(0.1\ \text{liter})}{(0.25\ \text{mole/liter})}$$

$$= \frac{10\ \text{g}/\cancel{\text{liter}}}{0.25\ \text{mole}/\cancel{\text{liter}}} = 40\ \text{g/mole}$$

MATERIALS:

- beaker, 250-ml
- sodium hydroxide (NaOH) solution, 0.1 M
- buret, 50-ml
- ring stand
- ring-stand clamp
- 2 beakers, 150-ml
- hydrochloric acid (HCl) solution, unknown concentration
- beaker, 50-ml
- pipet, 10-ml
- 2 Erlenmeyer flasks, 125-ml
- medicine dropper
- phenol red indicator solution
- vinegar (acetic acid--$HC_2H_3O_2$) solution, unknown concentration
- unknown solid acid
- balance
- graduated cyliner, 10-ml

PROCEDURE:

PART I: CARE AND USE OF THE BURET

1. Obtain approximately 200 ml of sodium hydroxide solution in a 250-ml beaker.

2. After making sure that the buret is clean, rinse it with about 10 to 20 ml of the NaOH solution. Take care not to spill any NaOH solution. Allow a few ml of the solution to flow through the tip, and then turn the buret horizontally so that the inside is completely wetted. Pour the remaining solution out of the top of the buret. A beaker (150-ml) may be used to collect waste NaOH solution that is not used in the titration.

3. Clamp the buret to a ring stand. Carefully pour NaOH solution into the buret until it is filled to near the top.

4. Dislodge any bubbles in the tip by allowing a small amount of the solution to flow rapidly out of the buret.

5. Note the markings on the buret. The graduations are given for each 0.1 ml and typically run from top to bottom (0 to 50 ml). Volumes may be estimated to ± 0.05 ml. Remember to keep the meniscus at eye level when making volume readings.

6. During each titration, the NaOH solution is delivered from the buret until the endpoint is reached. The volume of solution used is determined by subtracting the initial reading from the final reading. For example, if the meniscus is initially at a level of 10.0 ml and drops to 22.6 ml, the volume of solution delivered is 22.6 - 10.0 = 12.6 ml.

7. Adjust the meniscus to any desired level near the top of the buret. If a drop is hanging from the tip, it should be discarded before the titration begins. You can do this by touching the tip to the side of the waste-collecting beaker. (Important: a drop remaining on the tip during the titration should be added to the reacting mixture.)

8. To increase the precision of your results, use at least 10.0 ml of NaOH solution in each titration. For example, if a titration required only 3 ml of NaOH, one could repeat the titration, starting with four or five times as much acid solution.

9. When the activity is completed, rinse the buret thoroughly with water before putting it away.

## PART II: TITRATION OF ACIDS OF UNKNOWN CONCENTRATION

1. Obtain about 15 ml of the unknown HCl solution in a small (50-ml) beaker.

2. With a 10-ml pipet, add 10.0 ml of the acid solution to a 125-ml flask.

3. Add about 10 ml of tap water to the flask. (Note that adding water to the flask does not change the number of moles of acid originally present in the 10-ml sample. Therefore, small amounts of water may be used as necessary to rinse the sides of the flask. This will not affect the acid sample volume that is used in the final calculations.)

4. With a medicine dropper, add 3 to 5 drops of phenol red solution to the flask.

5. Record the initial level of the buret meniscus.

6. Add NaOH solution from the buret into the flask. Swirl the contents of the flask continuously. A red color will begin to appear, which at first disappears quickly. When the red color starts to disappear slowly, this indicates that the titration is nearing completion. (It is often helpful to place a sheet of white paper under the flask to make the colors more visible.)

7. At this stage, slowly add more solution a drop at a time until a faint red color persists. This marks the endpoint of the titration. Be careful not to add too much NaOH solution, since this will make the results inaccurate.

8. Record the final level of the meniscus. Determine the total volume of NaOH solution used to the nearest 0.1 ml. Record this value in your data sheet.

9. If time allows, repeat the titration until you feel that precise results have been obtained.

10. Calculate the concentration of acid present in the initial 10-ml sample using the formula provided in the introduction. Record the result to the nearest 0.001 mole per liter.

11. Rinse thoroughly the pipet, flask and 50-ml beaker.

12. Prepare a diluted solution of vinegar by combining 10.0 ml vinegar with 40 ml tap water.

13. Using the diluted vinegar solution, repeat Steps 2 through 10. (Do not use the dry flask. It is needed for Part III.)

14. Since the vinegar was diluted by a factor of five, the concentration of acid present in the vinegar may be calculated by multiplying the concentration of the diluted solution by five. Record the result to the nearest 0.001 mole per liter.

15. Rinse thoroughly all glassware that contained vinegar.

## PART III: DETERMINING THE MOLECULAR WEIGHT OF AN UNKNOWN SOLID ACID

1. Obtain a few grams (less than a teaspoonful) of the unknown acid.

2. Weigh out 3.00 ± .01 g of the unknown.

3. Dissolve this quantity in about 60 ml of tap water in the 150-ml beaker. Pour the solution into the 100-ml graduate.

4. Add 10 to 20 ml of water to the beaker and swirl briefly in order to mix any remaining acid solution. Add this solution to the contents of the graduate.

5. Fill the graduate with water to the 100-ml mark. The concentration of the solution is now 3.00 g/0.100 liter, which is equal to 30.0 g/liter. Pour the contents of the flask into the clean, dry flask.

6. Perform Steps 2 through 10 of Part II to determine the concentration of this solution in moles per liter.

7. Calculate the molecular weight of the unknown acid by dividing the mass concentration of the unknown (30.0 g/liter) by the molarity of the unknown, as explained in the introduction. Record the result.

## DISCUSSION QUESTIONS:

1. Why is it important to rinse the buret with NaOH solution before beginning the titration?

2. Suppose that each molecule of an acid released two hydrogen ions into solution instead of just one. How would this affect the calculation of the molarity of the acid? of the hydrogen ion?

3. What are some of the sources of error in this activity? Which do you think are the most important?

## SAMPLE DATA SHEET:

final buret reading __________ ml

initial buret reading - __________ ml

volume of NaOH solution ($V_B$) __________ ml

$$\text{concentration of acid } (C_A) = \frac{C_BV_B}{V_A} = \frac{(0.1 \text{ mole/liter})V_B \cancel{\text{ml}}}{(10 \cancel{\text{ml}})}$$

$$= (0.01)(V_B) \text{ mole/liter}$$

$$= \_\_\_\_\_\_\_\_\_\_ \text{ M}$$

$$\text{molecular weight of unknown acid} = \frac{\text{mass concentration}}{\text{molar concentration}}$$

$$= \frac{30 \text{ g/}\cancel{\text{liter}}}{C_A \text{ moles/}\cancel{\text{liter}}}$$

$$= \_\_\_\_\_\_\_\_\_\_ \text{ g/mole}$$

## LABORATORY ACTIVITY 30:

## THE MEASUREMENT OF pH

## GENERAL INTRODUCTION:

The relationship of pH to acidity and basicity has already been discussed in the Student Text. The term, pH, is used very frequently in health and medicine. A small change in blood pH can be lethal. In the laboratory, an error in preparing a solution amounting to a tenth of a pH unit can lead to completely erroneous results.

Recall that the pH scale tells us how acidic or basic a substance is. It is based upon the concentration of hydrogen ions in a solution. The midpoint of the scale is seven, at which point the concentration of hydrogen ions ($H^+$) equals the concentration of hydroxide ions ($OH^-$). Solutions with a pH lower than seven are acidic. The lower the pH, the more acidic the solution. Solutions with a pH greater than seven are basic. The stronger the base, the higher the pH.

The pH of a solution can be measured by several methods. Two of these will be used in this activity. The first method involves chemical indicators which change colors at different pH values. You may have noticed that when red cabbage is boiled, the water takes on a red color. If a base is then added to the cabbage water, the fluid will change from red to green. This is because cabbage contains a pH indicator which changes color at a certain pH.

There are many different chemical pH indicators and they change colors at different points on the pH scale. So the color of a solution containing a pH indicator tells us the approximate pH. For example, litmus changes from blue to red at pH 7. A pH test paper can be made by soaking filter paper in litmus solution. Red litmus paper means that the solution must be acidic. Conversely, blue paper indicates a basic solution. It is possible to make a pH paper containing a number of indicators that changes color several times as the pH is changed--this permits a more accurate estimation of the pH of a solution.

Special indicators, such as Hydrion paper, show varying shades of color which correspond to the whole range of pH values. To measure the acidity or basicity of a solution, one wets the Hydrion paper with the solution and locates the resulting color shade on the standard color chart found on the side of the Hydrion paper container. The pH is listed above each color shade on the standard color chart.

In this activity you will use Hydrion paper to determine the pH value of various solutions. However, in the clinical (hospital) laboratory, pH is generally determined with a pH meter rather than with pH paper. This is because the pH meter can provide much more accurate measurements of pH than pH paper. The BIP pH meter which you will use in Part III may be assumed to be accurate to the nearest 0.1 to 0.2 pH units. The pH meter measures the electric current passing through a probe that is dipped into a solution.

The probe you will use is a combination electrode. The active component of the probe is a "glass electrode." This electrode holds $H^+$ ions at a known concentration (0.1 M HCl). When this electrode is immersed in a second solution of a different hydrogen ion concentration, $H^+$ ions will tend to flow from one solution to the other. This creates a voltage drop and a flow of electrons which is registered by the instrument as pH.

## PART I: pH OF A VARIETY OF SOLUTIONS

### INTRODUCTION:

In the first part of the activity, the pH of a variety of solutions will be determined with Hydrion paper.

### MATERIALS:

one roll each of Hydrion pH test paper (1-11 range) and Hydrion short-range pH test paper (1.0-5.5 range) and respective color code charts

forceps

set of test solutions in beakers or other containers

### PROCEDURE:

1. Tear a 1-cm length of Hydrion pH paper (1-11 range) from the roll. With forceps immerse about half of the paper into the solution to be tested. Remove the Hydrion paper and quickly check its color against the color code chart. This must be done quickly because the color changes as the paper dries out. Record the pH in the data table.

2. Test each solution in the following table with a fresh strip of this Hydrion paper (pH 1-11). If you accidentally dip the forceps into a solution, rinse them off and wipe them with a paper towel. Record your observations. Note the pH range of the last two solutions is not given--these are unknowns for you to figure out.

3. If you have any solutions in the pH range of 1-5.5, check their pH with the 1-5.5 pH paper, too.

| Solution | Expected pH Range |
|---|---|
| Tap Water | 4.0 - 8.0 |
| Distilled Water | 5.0 - 7.0 |
| Vinegar | 2.4 - 3.4 |
| Household Ammonia | 10.0 - 12.0 |
| Saliva | (not given) |
| Urine | (not given) |

## PART II: EFFECT OF DILUTION ON pH

### INTRODUCTION:

In laboratory investigations, acids of known pH must often be prepared. Generally this is done by diluting a concentrated acid solution with water. In this part of the activity, you will dilute two different acids (hydrochloric acid and acetic acid) and compare the resulting pH values.

If 1 ml of a $10^{-1}$ molar acid solution is transferred to a second test tube and 9 ml of water is added, then the second solution will be one-tenth as concentrated as the first--the concentration will be $10^{-2}$ molar. The dilution process may be repeated again--1 ml from the second tube may be transferred to a third tube and 9 ml of water may be added to this tube. The concentration of acid in the third tube would then be $10^{-3}$ molar. If the process is repeated again, the resulting concentration will be $10^{-4}$ molar. The dilution process is called "serial dilution." In Part II, you will do serial dilutions of hydrochloric acid (HCl) and acetic acid. The object will be to determine the effect of dilution on pH and to compare the pH of two different acids at identical concentrations.

### MATERIALS:

one roll of Hydrion test paper (1.0 to 5.5 range) and color code chart
forceps
8 test tubes (16 x 125 mm)
test-tube rack
2 pipets, 10-ml
20 ml of $10^{-1}$ M hydrochloric acid
20 ml of $10^{-1}$ M acetic acid

PROCEDURE:

1. Prepare dilutions of HCl of $10^{-2}$, $10^{-3}$ and $10^{-4}$ M from a $10^{-1}$ M stock solution by serial dilutions as reviewed in the figure.

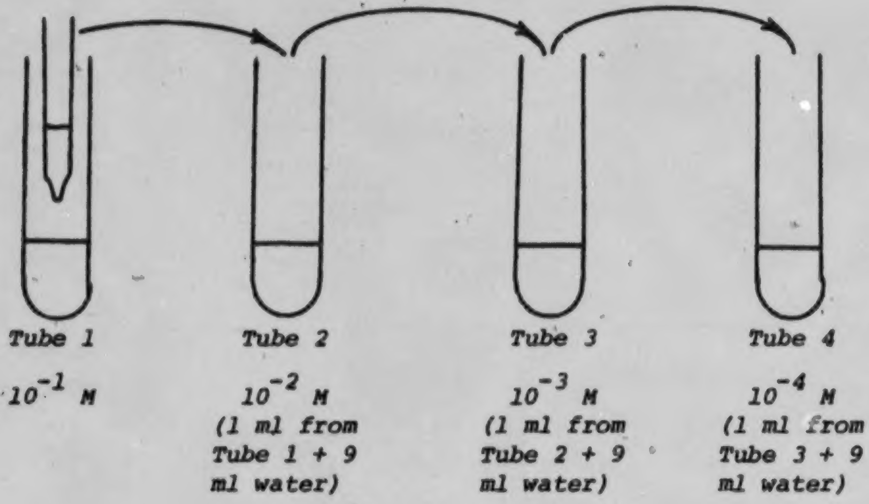

Tube 1
$10^{-1}$ M
Tube 2
$10^{-2}$ M
(1 ml from Tube 1 + 9 ml water)
Tube 3
$10^{-3}$ M
(1 ml from Tube 2 + 9 ml water)
Tube 4
$10^{-4}$ M
(1 ml from Tube 3 + 9 ml water)

FIGURE 1: *Serial dilution.*

Notes: A. HCl is toxic and should not be pipetted by mouth. Many other acids are toxic too; so, unless you are told otherwise, do not pipet an acid by mouth.

B. After you add 1 ml of $10^{-1}$ M HCl and 9 ml water to Tube 2, mix the contents very thoroughly. Otherwise there will be a large error in preparing Tubes 3 and 4. The mixing procedure must be done again after Tube 3 is prepared and after Tube 4 is prepared.

C. Be sure to label all tubes.

2. Determine the pH of each solution with the narrow range pH paper and record the values.

3. Repeat Steps 1 and 2 with $10^{-1}$ acetic acid replacing the $10^{-1}$ M HCl.

4. Determine and record the pH of each solution with the pH paper in the range of 1-5.5.

5. Convert molar concentration to p(concentration) and then graph pH as a function of p(concentration) for the two acids. Use pH for the vertical coordinates and show each pH value as a range of imprecision. You may need to review Section 30 of the Student Text. After you complete the graph, write an equation relating pH to p(concentration) for each acid.

## PART III: USE OF A pH METER

### INTRODUCTION:

The BIP pH meter will be introduced briefly in this activity and then will be used more extensively in later activities. The object will be to learn how to use the instrument and also to compare readings of pH made with indicators and with a pH meter.

pH Probe Cautions:

1. The glass probe and especially the bulb at the bottom of the probe is very delicate. Do not touch it. It is given some protection by the plastic housing.

2. When not in use, the probe should be left immersed in distilled water. A probe that dries out may become very sluggish in response and may require a time-consuming readjustment of the instrument.

### MATERIALS:

BIP
small screwdriver
pH probe
distilled water

4 beakers, 50-ml
pH 7 solution
urine
small ring stand and clamp

### PROCEDURE:

1. In order to use the BIP to measure pH, some adjustments have to be made. Before continuing, be sure that you have read the "pH Probe Cautions" described in the introduction. The pH probe should be supported as in Figure 2. Otherwise, the probe may cause the beaker to topple over and the delicate glass probe would be likely to break.

2. Check to see whether the following has been done. The probe has been supported as in Figure 2. The probe has been plugged into the BIP in the position labeled 1 in Figure 3 on the following page. The BIP is programmed: F to G, H to K. The BIP has been plugged into a current source.

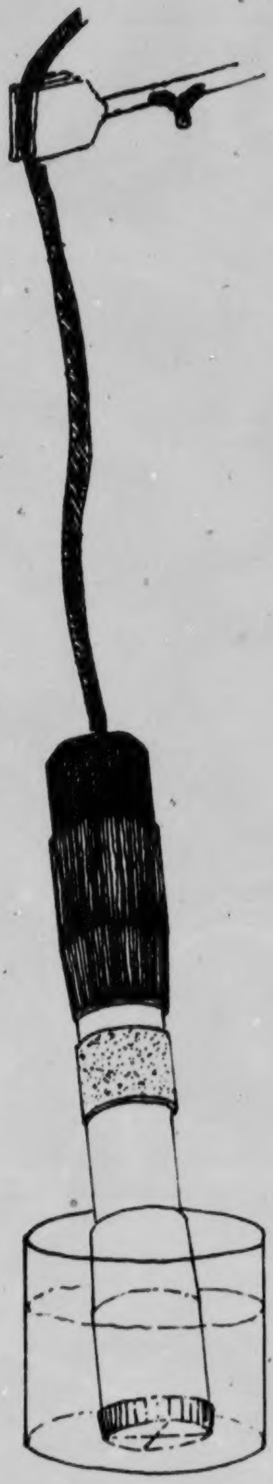


FIGURE 2

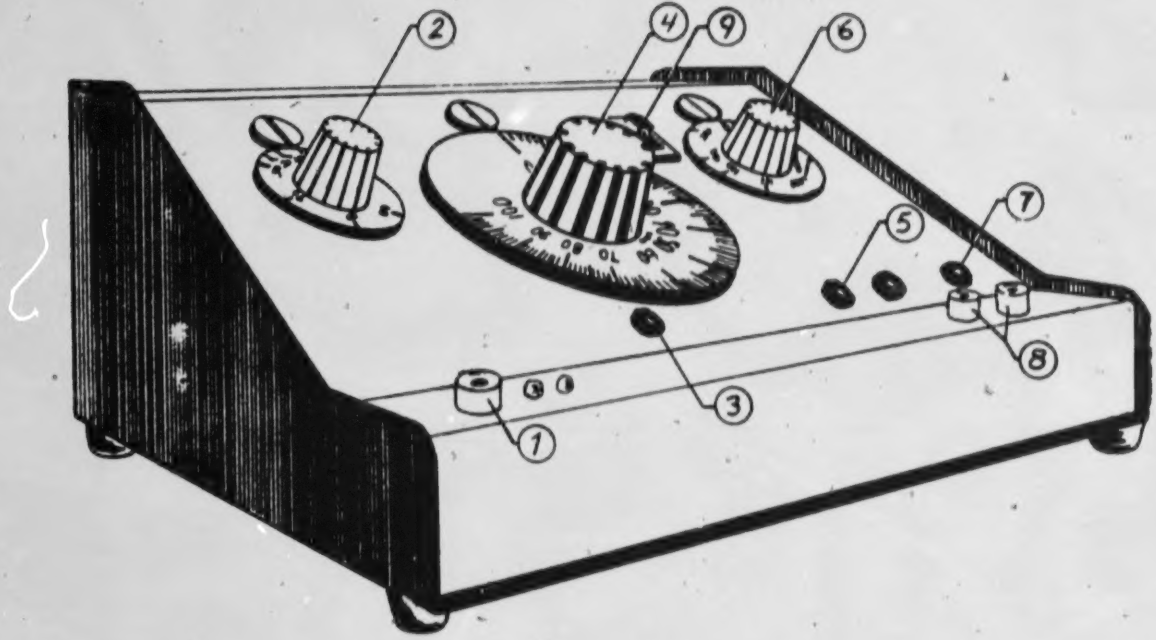


FIGURE 3: *Front view of the BIP.*

3. The pH meter must be standardized (pre-set) with a solution of known pH. To do this, remove the cap from the pH probe with care.

4. Rinse the probe thoroughly by gentle agitation in about 25 ml of distilled water in a 50-ml beaker. It is desirable to support the probe with a buret clamp on a small ring stand as in Figure 2.

5. Before determining the pH of an unknown, the BIP pH meter must be standardized against a solution of known pH. In this case, use a solution of pH 7.

6. Let the water drain off the surface of the probe, without actually touching the glass part of the probe. Immerse the probe in 25 ml of pH 7 solution contained in a 50-ml beaker. Gently agitate for 10 to 15 seconds.

7. Set the mA dial at 70 which corresponds to a pH of 7.0 ($\text{pH} = \frac{\text{dial reading}}{10}$).

8. Turn the slotted control (Figure 3, Position 5) with a screwdriver until the diode lights indicate a null condition. This adjustment should be made after about 30 seconds contact of the probe with the pH 7 solution. This adjustment need be done only once at the beginning of a lab if it is desired to determine the pH of a series of unknowns.

9. After standardization (Steps 5 through 8 above), rinse the probe thoroughly with at least two changes of distilled water. The same beaker may be used for each rinse.

10. Pour about 25 ml of the urine solution into a 50-ml beaker. This solution was picked because it is usually close to pH 7 and the pH meter was standardized at that pH. pH meters are most accurate with solutions at pH values close to the pH used for standardization.

11. Immerse the probe in the urine solution and gently agitate for 10 to 15 seconds. After another 15 seconds, turn the mA dial to null. Record the BIP number and the pH. Compare it with the reading made with indicator paper.

SAMPLE DATA TABLES:

PART I: pH OF A VARIETY OF SOLUTIONS

| Solution | Expected pH Range | pH Obtained |
|---|---|---|
| Tap Water | 4.0 - 8.0 | |
| Distilled Water | 6.0 - 7.0 | |
| | | |

PART II: EFFECT OF DILUTION ON pH

| Hydrochloric Acid | | | | Acetic Acid | | | |
|---|---|---|---|---|---|---|---|
| Tube | Molar Conc. | p(concentration) | pH | Tube | Molar Conc. | p(concentration) | pH |
| 1 | $10^{-1}$ | | | 1 | $10^{-1}$ | | |
| 2 | $10^{-2}$ | | | 2 | $10^{-2}$ | | |
| | | | | | | | |

PART III: USE OF A pH METER

| Solution | pH Obtained With pH Paper (Part I) | pH Obtained With pH Meter |
|---|---|---|
| Urine | | |
| | | |

BIP # ________

DISCUSSION QUESTIONS:

PART I:

1. Why is it important to measure the pH of body fluids?

2. How did your pH measurements compare with the expected results? Can you think of any reasons why your values may not have agreed with the expected results? (Can you think of any possible sources of error?)

3. Which substances tested were acids? Which substances tested were bases? Explain.

4. How precisely can you measure pH with each type of pH paper?

5. Compare your results with other students in your class. What reasons can you give for any differences between the results?

PART II:

1. Which of the acids tested in Part II was the stronger acid? Explain.

2. How is it that two acids at the same concentration can have different pH values?

3. How might your results be different if you had forgotten to mix Tube 2 after adding 1 ml from Tube 1 and 9 ml water?

4. In Part II, there is a warning about the dangers of working with acids. Yet acetic acid (vinegar) is sometimes used in making salads. How is this possible?

5. Compare the relation of pH to concentration for HCl and acetic acid.

PART III:

1. Compare the precision of the two methods you have used for measurement of pH.

2. Suppose you had pH paper and a pH meter. (a) Can you think of a situation in which you might prefer to measure the pH of a solution with the pH meter? (b) Can you also think of an opposite example--when you would prefer to use the pH paper?

## LABORATORY ACTIVITY 31:

## COLORIMETRY

### INTRODUCTION:

Many substances, when dissolved in water, produce a colored solution. One example is copper sulfate, which you used for electroplating. If we add just a tiny amount of copper sulfate to water, the resulting solution takes on a very pale tinge of blue. But as we add more and more $CuSO_4$ crystals, the solution takes on a progressively deeper and more intense blue color.

It should come as no surprise that the depth and intensity of the color is related to the concentration of the dissolved substance imparting that color. The relation between color and concentration is by no means a simple one, but it is well understood.

If we shine a light through a colorless liquid, such as pure water, part of the light is absorbed by the liquid and part is transmitted through the liquid. The transmitted light--the light that comes out on the other side--is always less intense than the light that entered the liquid to start with. This is true because some of the light is always absorbed (see Figure 1).

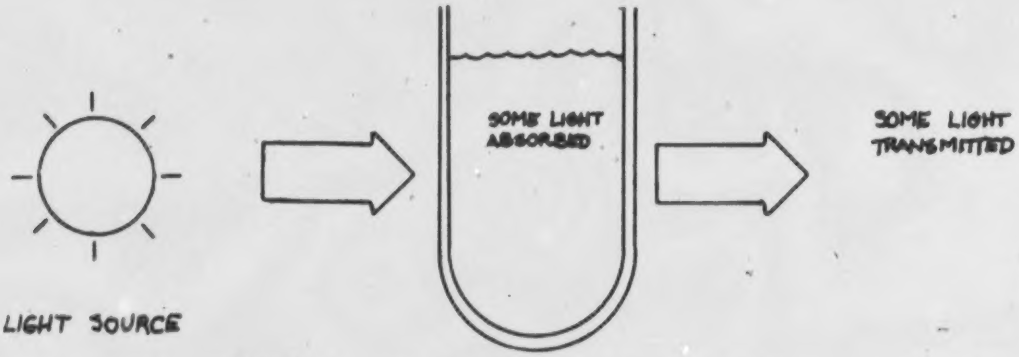
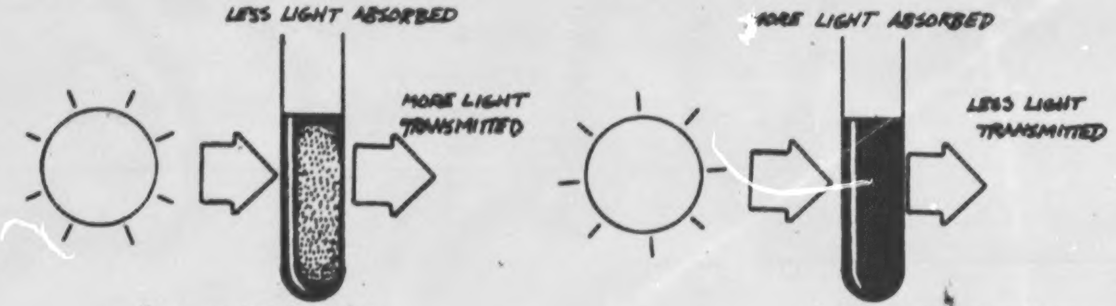


FIGURE 1: *Absorbance and transmittance of light by a liquid.*

When a colored substance is dissolved in water, the colored solution absorbs more light than pure water does. Therefore less light is transmitted. If we increase the concentration, adding more color to the solution, still more light is absorbed and still less is transmitted (see Figure 2 on the next page).

If we had a way to measure the proportion of light transmitted by solutions of different known concentrations, we could estimate the concentrations of unknown solutions by the same process. This is one of the many things that the BIP--with one attachment--can do. The process is called colorimetry (KUH-luh-RIM-uh-tree). It is a technique that is very commonly used in the biomedical analysis of body fluids. In fact, a large portion of all medical laboratory procedures depend upon colorimetric analysis.



*low concentration --pale color*

*high concentration-- deep, intense color*

FIGURE 2: *The relation of concentration to absorbance and transmittance of light.*

In order to make colorimetric measurements, the BIP is connected to a colorimeter test well. As shown in Figure 3, the test well contains a circular opening which is just the right size to hold a test tube that is 16 mm in diameter. A tube used for colorimetry is often referred to as a cuvet (kew-VET).

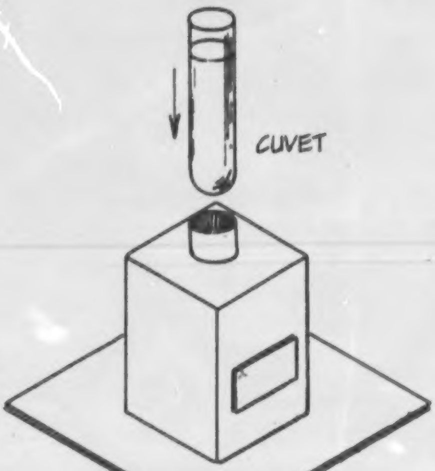


FIGURE 3: *The colorimeter test well.*

The inside of the test well consists basically of just two things. On one side of the well is a light source. On the opposite side is a photocell. The test well is connected to the BIP, and a cuvet containing a solution is placed in the well. The light source shines light through the solution, and the photocell detects the percentage of that light that is transmitted through the solution.

This quantity, which can be read off the mA dial after suitable adjustment of the BIP, is referred to as percentage transmittance and is often abbreviated % T.

The quantity of light absorbed by the solution, the absorbance, may be determined from a table of values that relates % T to absorbance.

The purpose of this activity is to familiarize you with the techniques of colorimetry and the relation between concentration, absorbance and transmittance. Pay attention, because you will need to understand these procedures in order to perform many of the activities to come.

MATERIALS:

5.6 g chromium potassium sulfate, $CrK(SO_4)_2\cdot 12H_2O$
graduated cylinder, 100-ml
2 beakers, 150-ml (or larger)
2 pipets, 10-ml
test tube, 16 x 125 mm
test-tube rack
balance
mortar and pestle (optional)
BIP and colorimeter test well
programming wire
wire cutter-strippers
screwdriver
Kimwipes or tissue
parafilm

PROCEDURE:

1. On the side of the test well you will find three connecting pins, each of which is identified by a letter. The letter indicates that the pin is to be connected to the pin with that same letter on the programming panel of the BIP. For example, the pin marked "S" on the test well is to be connected to pin S on the BIP.

2. Use programming wire to make the three connections indicated between the test well and the BIP.

3. In addition, make the following connections on the BIP programming panel: H to L and T to U.

4. Check your programming to make sure it is correct. Only then plug in the BIP. Allow at least ten minutes of warm-up time--the longer the better. While you are waiting, complete Steps 5 through 8.

5. Weigh out 5.6 ± .05 g of chromium potassium sulfate [$CrK(SO_4)_2\cdot 12H_2O$].

6. If you have crystals, grind them to a powder with a mortar and pestle before doing the next step. Large $CrK(SO_4)_2\cdot 12H_2O$ crystals dissolve very slowly in water.

7. Put the $CrK(SO_4)_2\cdot 12H_2O$ into a 100-ml graduated cylinder and add enough tap water to make 100 ± 1 ml of solution.

8. Pour the solution back and forth between the graduated cylinder and a 150-ml (or larger) beaker until all of the $CrK(SO_4)_2 \cdot 12H_2O$ is dissolved. Store it in the beaker for use later on.

9. If the BIP is warmed up, it is time to standardize the colorimeter. What you are going to be doing later is determining the % T for a series of known concentrations of chromium potassium sulfate solution. Both the water and the walls of the cuvet absorb some light, but this problem can be compensated for by an adjustment of the BIP. The BIP can be adjusted so that it gives a reading of 100% T (in other words, no light absorbed) for a cuvet containing only water. Once this adjustment has been made, the effect of the solvent (tap water, in this case) and the cuvet have been compensated for. This process of adjustment is called standardization of the colorimeter, and the cuvet containing only solvent is often referred to as a blank.

10. Add 10 ± 1 ml of tap water to a cuvet. This is the blank.

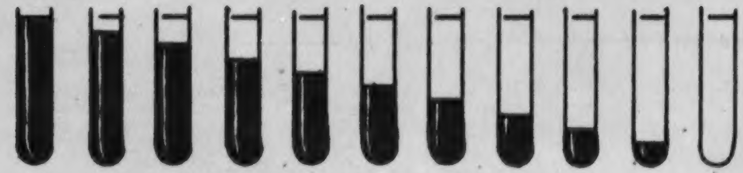
11. Use a tissue to wipe the outside of the cuvet free of liquid, fingerprints, etc. This step is very important for two reasons. If the tube is not clean (both inside and out), the amount of light transmitted will be affected. In addition, liquid on the outside of the cuvet could easily damage the components of the test well.

12. Insert the cuvet into the test well. Use only 16-mm test tubes in the test well. If a test tube must be forced into the well, it is too large and should be discarded.

13. With the BIP programmed for colorimetry, the numbers on the mA dial give the percentage transmittance. Carefully set the mA dial on 100% T.

14. Insert a screwdriver into the slotted control of the voltage regulator, which is just below and to the left of the mA dial. Adjust the control to the null point. This is the point at which the slightest turn in either direction will change which of the two diode lights is on. Remove the screwdriver. The colorimeter is now standardized. Remove the cuvet from the test well and empty it as completely as possible.

In the remaining steps you will be measuring the % T of the 11 solutions shown in Figure 4 on the next page. The "indicator" referred to is the chromium potassium sulfate solution you prepared in Steps 5 through 8.



| | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|
| ml of indicator | 10 | 9 | 8 | 7 | 6 | 5 | 4 | 3 | 2 | 1 | 0 |
| ml of tap water | 0 | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 |
| total ml | 10 | 10 | 10 | 10 | 10 | 10 | 10 | 10 | 10 | 10 | 10 |

FIGURE 4: *Directions for preparing eleven concentrations of the indicator solution.*

Each volume of indicator and each volume of tap water should be measured to the nearest 0.1 ml. Remember to use separate pipets for each of the two liquids. We will describe in detail the steps to be taken for the first two concentrations.

15. Add 10.0 ml of indicator to the cuvet. Wipe the cuvet and place it in the test well.

16. Turn the mA dial to the null point. Record the % T to the nearest 0.5%.

17. Remove the cuvet from the test well and dispose of its contents. Drain the cuvet as completely as possible.

18. Add 9.0 ml of the indicator and 1.0 ml of tap water to the cuvet. To mix the solution cover the mouth of the cuvet with parafilm. (To make a watertight seal, dry the outside of the cuvet if it is wet. Then gently stretch the parafilm over the mouth of the cuvet. It will stick tightly to the glass.) Place your thumb over the parafilm and invert the cuvet three or four times to mix the solution thoroughly.

19. Remove the parafilm, wipe the cuvet and insert it into the test well.

20. Determine the % T and record it.

21. Repeat Steps 17 through 20 for each of the remaining concentrations shown in Figure 4.

22. Use the Absorbance-Transmittance Table to record the absorbance for each % T that you have recorded.

23. Graph percentage transmittance as a function of concentration. Use % T for the vertical coordinates and ml indicator per 10 ml of solution as the horizontal coordinates.

24. On a second piece of graph paper, graph absorbance as a function of concentration. Use absorbance as the vertical coordinates. Draw the "best" straight line through your data points. [Note: Absorbances greater than about 0.6 are likely to be more inaccurate than absorbances between 0 and 0.6. It may be a good idea to ignore the data points with absorbances greater than 0.6 in drawing a "best" line. In addition, the "best" line should pass through or near the origin (0,0).]

SAMPLE DATA SHEET:

(Always record the number of the BIP and test well used to obtain the data. That way if the data ever suggest that the equipment was malfunctioning, it will be possible to identify the instrument in question.)

BIP # ________ Test Well # ________

| ml Indicator per 10 ml of solution | % T | Absorbance |
|---|---|---|
| 10 | | |
| 9 | | |
| 8 | | |

DISCUSSION QUESTIONS:

1. As the concentration increases, does the percentage transmittance increase or decrease? Does the absorbance increase or decrease?

2. Why is it necessary to use a blank to standardize the colorimeter?

3. What are the main components inside the colorimeter test well? What is the function of each?

4. Give two reasons for always wiping a cuvet clean before inserting it into the test well.

5. Why is it important to mix a solution thoroughly before measuring its transmittance?

6. This activity stops one step short of the primary function of colorimetry, which is to determine the concentration of unknown solutions. Assume that you have tested a solution containing an unknown number of ml of the indicator in 10 ml of solution. The % T is measured to be 42.5.

a. Use your graph of % T as a function of concentration to estimate the concentration of the unknown.

b. Use your graph of absorbance as a function of concentration to estimate the concentration of the unknown.

c. Do you think one of these two estimates is better than the other? Explain your answer.

d. What if you had only one or two known data points on each graph? Could you still give a reasonable estimate of the unknown concentration? In this case, would one graph give a more reliable estimate than the other? Explain your answer.

# ABSORBANCE-TRANSMITTANCE TABLE

| Trans. (%) | Absorbance | Trans. (%) | Absorbance | Trans. (%) | Absorbance | Trans. (%) | Absorbance |
|---|---|---|---|---|---|---|---|
| 0.0 | ----- | 25.0 | .600 | 50.0 | .300 | 75.0 | .125 |
| 0.5 | 2.300 | 25.5 | .595 | 50.5 | .295 | 75.5 | .120 |
| 1.0 | 2.000 | 26.0 | .585 | 51.0 | .290 | 76.0 | .120 |
| 1.5 | 1.825 | 26.5 | .575 | 51.5 | .290 | 76.5 | .115 |
| 2.0 | 1.700 | 27.0 | .570 | 52.0 | .285 | 77.0 | .115 |
| 2.5 | 1.600 | 27.5 | .560 | 52.5 | .280 | 77.5 | .110 |
| 3.0 | 1.525 | 28.0 | .555 | 53.0 | .275 | 78.0 | .110 |
| 3.5 | 1.460 | 28.5 | .545 | 53.5 | .270 | 78.5 | .105 |
| 4.0 | 1.400 | 29.0 | .540 | 54.0 | .270 | 79.0 | .100 |
| 4.5 | 1.345 | 29.5 | .530 | 54.5 | .265 | 79.5 | .100 |
| 5.0 | 1.300 | 30.0 | .525 | 55.0 | .260 | 80.0 | .095 |
| 5.5 | 1.260 | 30.5 | .515 | 55.5 | .255 | 80.5 | .095 |
| 6.0 | 1.220 | 31.0 | .510 | 56.0 | .250 | 81.0 | .090 |
| 6.5 | 1.185 | 31.5 | .500 | 56.5 | .250 | 81.5 | .090 |
| 7.0 | 1.155 | 32.0 | .495 | 57.0 | .245 | 82.0 | .085 |
| 7.5 | 1.125 | 32.5 | .490 | 57.5 | .240 | 82.5 | .085 |
| 8.0 | 1.095 | 33.0 | .480 | 58.0 | .235 | 83.0 | .080 |
| 8.5 | 1.070 | 33.5 | .475 | 58.5 | .230 | 83.5 | .080 |
| 9.0 | 1.050 | 34.0 | .470 | 59.0 | .230 | 84.0 | .075 |
| 9.5 | 1.020 | 34.5 | .460 | 59.5 | .225 | 84.5 | .075 |
| 10.0 | 1.000 | 35.0 | .455 | 60.0 | .220 | 85.0 | .070 |
| 10.5 | .980 | 35.5 | .450 | 60.5 | .220 | 85.5 | .070 |
| 11.0 | .960 | 36.0 | .445 | 61.0 | .215 | 86.0 | .065 |
| 11.5 | .940 | 36.5 | .440 | 61.5 | .210 | 86.5 | .065 |
| 12.0 | .920 | 37.0 | .430 | 62.0 | .210 | 87.0 | .060 |
| 12.5 | .905 | 37.5 | .425 | 62.5 | .205 | 87.5 | .060 |
| 13.0 | .885 | 38.0 | .420 | 63.0 | .200 | 88.0 | .055 |
| 13.5 | .870 | 38.5 | .415 | 63.5 | .200 | 88.5 | .055 |
| 14.0 | .855 | 39.0 | .410 | 64.0 | .195 | 89.0 | .050 |
| 14.5 | .840 | 39.5 | .405 | 64.5 | .190 | 89.5 | .050 |
| 15.0 | .825 | 40.0 | .400 | 65.0 | .185 | 90.0 | .050 |
| 15.5 | .810 | 40.5 | .395 | 65.5 | .185 | 90.5 | .045 |
| 16.0 | .795 | 41.0 | .390 | 66.0 | .180 | 91.0 | .040 |
| 16.5 | .785 | 41.5 | .380 | 66.5 | .175 | 91.5 | .040 |
| 17.0 | .770 | 42.0 | .375 | 67.0 | .175 | 92.0 | .035 |
| 17.5 | .755 | 42.5 | .370 | 67.5 | .170 | 92.5 | .035 |
| 18.0 | .745 | 43.0 | .365 | 68.0 | .165 | 93.0 | .030 |
| 18.5 | .735 | 43.5 | .360 | 68.5 | .165 | 93.5 | .030 |
| 19.0 | .720 | 44.0 | .355 | 69.0 | .160 | 94.0 | .025 |
| 19.5 | .710 | 44.5 | .350 | 69.5 | .160 | 94.5 | .025 |
| 20.0 | .700 | 45.0 | .345 | 70.0 | .155 | 95.0 | .020 |
| 20.5 | .690 | 45.5 | .340 | 70.5 | .150 | 95.5 | .020 |
| 21.0 | .680 | 46.0 | .335 | 71.0 | .150 | 96.0 | .015 |
| 21.5 | .670 | 46.5 | .330 | 71.5 | .145 | 96.5 | .015 |
| 22.0 | .660 | 47.0 | .330 | 72.0 | .145 | 97.0 | .015 |
| 22.5 | .650 | 47.5 | .325 | 72.5 | .140 | 97.5 | .010 |
| 23.0 | .640 | 48.0 | .320 | 73.0 | .135 | 98.0 | .010 |
| 23.5 | .625 | 48.5 | .315 | 73.5 | .135 | 98.5 | .005 |
| 24.0 | .620 | 49.0 | .310 | 74.0 | .130 | 99.0 | .005 |
| 24.5 | .610 | 49.5 | .305 | 74.5 | .130 | 99.5 | .000 |
|  |  |  |  |  |  | 100.0 | .000 |

## LABORATORY ACTIVITY 32:

## TESTING EXPIRED AIR FOR CONCENTRATION OF CARBON DIOXIDE

### INTRODUCTION:

You may recall in Laboratory Activity 9 testing your breath for the presence of carbon dioxide. The test used there was one in which expired air was mixed with phenol red solution. Phenol red is an indicator that is red in basic solutions and yellow in acidic ones. As the $CO_2$ from your breath dissolved in the solution, part of it combined with water to form carbonic acid, which in turn dissociated into $H^+$ and $HCO_3^-$ ions. The increase in the concentration of $H^+$ ions caused the phenol red to turn yellow. Notice that the phenol red test is not specific for $CO_2$. The same result could be obtained by adding any acid to phenol red.

In this activity you will again test your breath for $CO_2$. In this case, however, it will be possible to determine the percentage of $CO_2$ in your breath, both for a normal breath and after various periods of breath-holding.

The procedure involves a chemical reaction between the $CO_2$ expired and a solution of barium hydroxide.

$$CO_2 + Ba(OH)_2 \rightarrow BaCO_3 + H_2O$$

The barium combines with the carbon dioxide and one atom of oxygen to give barium carbonate. The $Ba(OH)_2$ is a clear, colorless solution. The $BaCO_3$ forms a white precipitate, giving a cloudy appearance to the solution. If allowed to stand, the $BaCO_3$ will settle to the bottom of the container.

The techniques used to determine the amount of $BaCO_3$ that is formed are very similar to the colorimetric techniques used in Laboratory Activity 31. In colorimetry the percentage of light transmitted depends upon the amount of light that is absorbed by colored substances that are dissolved in the solvent.

In this procedure, the $BaCO_3$ makes the solution cloudy or turbid. The greater the concentration of $BaCO_3$, the greater the turbidity of the solution and the more light is absorbed by the $BaCO_3$ particles. The process of measuring the percentage of light transmitted by turbid solutions is called turbidimetry.

There are only two important differences between colorimetry and turbidimetry as far as laboratory procedure goes. In turbidimetry, the particles are not dissolved and tend to sink toward the bottom of the cuvet. Therefore, it is important that the particles be uniformly mixed throughout the liquid just before a final reading is taken.

The other difference has to do with the range over which the graph of absorbance vs. concentration is linear. In colorimetry, we indicated that absorbance is a linear function of concentration for absorbances between 0 and about .6, which correspond to % T readings between 100 and about 25%. In turbidimetry, absorbance is a linear function of concentration only for absorbances between 0 and .3, which correspond to % T readings between 100 and 50%. In either case, it is always possible to make a dilution that will bring the readings into the desired range.

Normally, turbidimetry requires a determination of the absorbances of a set of standard solutions of known concentration, from which a standard graph can be drawn. The concentration of an unknown is then determined from the graph. In this case, however, we have provided a standard graph at the end of the procedure which relates absorbance directly to the percentage $CO_2$ in the expired air.

## MATERIALS:

BIP and colorimeter test well

70 ml $Ba(OH)_2$ solution--keep stoppered when not in use!

Erlenmeyer flask, 250-ml, with stopper

Erlenmeyer flask, 125-ml, with stopper

graduated cylinder, 100-ml

graduated cylinder, 10-ml

6 cuvets, 16 x 125 mm, with stopper

test-tube rack

glass-marking pencil

Kimwipes or other tissues

clock with second hand

## PROCEDURE:

1. Program the BIP and test well for colorimetry. (See Laboratory Activity 31 for the programming instructions.)

2. After you have checked the programming carefully, plug in the BIP. Allow at least 10 minutes warm-up time--the longer the better.

3. Obtain about 70 ml of $Ba(OH)_2$ solution in a 125-ml flask. Keep this solution stoppered whenever possible, because it will react with $CO_2$ from the air and throw off your results if left open too long. $Ba(OH)_2$ is also a fairly strong base. If you should get any of the solution on your skin, wash it off with plenty of water.

4. Add 10 ml of tap water to each of the six cuvets. Mark the 10-ml level with a glass-marking pencil. Number the cuvets from 1 through 6. Drain them as completely as possible, and set them aside.

5. Use a 10-ml graduated cylinder to add 10 ml of $Ba(OH)_2$ solution to cuvet #1. Stopper immediately. (Remember to keep all samples of the solution stoppered when not in use.) This will be the blank.

6. Add 10 ml of $Ba(OH)_2$ solution to a 250-ml flask. Stopper immediately. Try to keep the liuqid from wetting the lip of the flask. If the lip does get wet, wipe it dry with a tissue.

7. Read all of this step before proceeding. Take a deep breath, remove the stopper and exhale as completely as possible into the flask. Allow air to excape from the flask as you blow in. Do not put your lip on the lip of the flask. At the end of the breath, quickly stopper the flask again.

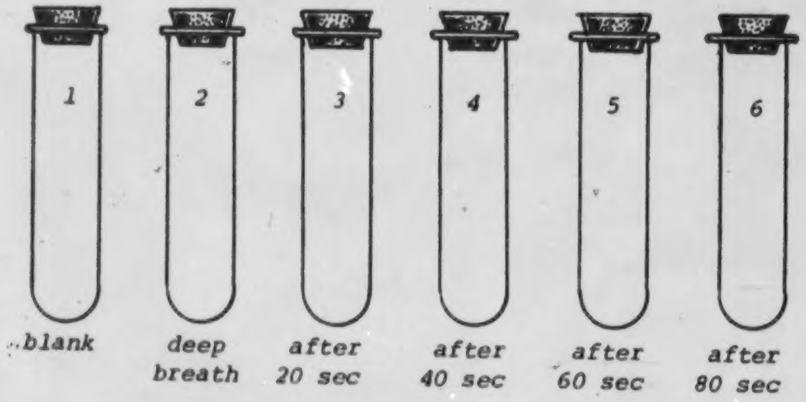
8. Shake the flask vigorously for 15 seconds. This will permit the $Ba(OH)_2$ to come into contact with most of the $CO_2$ in the flask.

9. Add 90 ml of tap water to the flask. Re-stopper and shake the flask to mix the contents thoroughly. This dilution is necessary to bring the concentration of $BaCO_3$ into the range of linearity discussed in the introduction.

10. As soon as you are through shaking, quickly pour 10 ml of the diluted solution into cuvet #2. Stopper the cuvet and put it aside. The transfer should be make quickly so that the $BaCO_3$ doesn't have time to settle toward the bottom of the flask.

11. Wash, rinse and drain the 250-ml flask.

12. Repeat Steps 6 through 11 four more times. The only difference is to hold your breath for the time periods shown in the following diagram before exhaling into the flask.



13. Wipe the blank (cuvet #1) and place it in the test well. Adjust the BIP so that the diode lights null at 100% transmittance. (See Laboaratory Activity 31 for the procedure.)

14. Wipe off cuvet #2. Invert it to mix the solid particles of $BaCO_3$ throughout the solution, then place it in the test well. Turn the mA dial until the light diodes null.

15. While you were adjusting the mA dial, some of the solid particles were settling to the bottom of the cuvet, causing your reading to be off slightly. Remove the cuvet, invert it once again, quickly return it to the test well and adjust the mA dial. The second reading should require only a slight change in the dial and should be made quickly. Record this reading on your data sheet.

16. Repeat Steps 15 and 16 for cuvets #3 through #6.

17. After you have cleaned up all equipment, use the table in Laboratory Activity 31 to convert your data to absorbances.

18. Use the graph to determine and record the percentage $CO_2$ in each of the five breath samples.

DISCUSSION QUESTIONS:

1. Under what conditions did your breath contain the greatest concentration of $CO_2$?

2. Under what conditions do you think even more $CO_2$ would be present in the air you exhale?

3. Under what condition might your breath contain less than a normal concentration of $CO_2$?

4. What effect does holding your breath have on the concentration of $CO_2$ in your blood?

5. What has happened in your body when you hold your breath until you involuntarily exhale and begin breathing again?

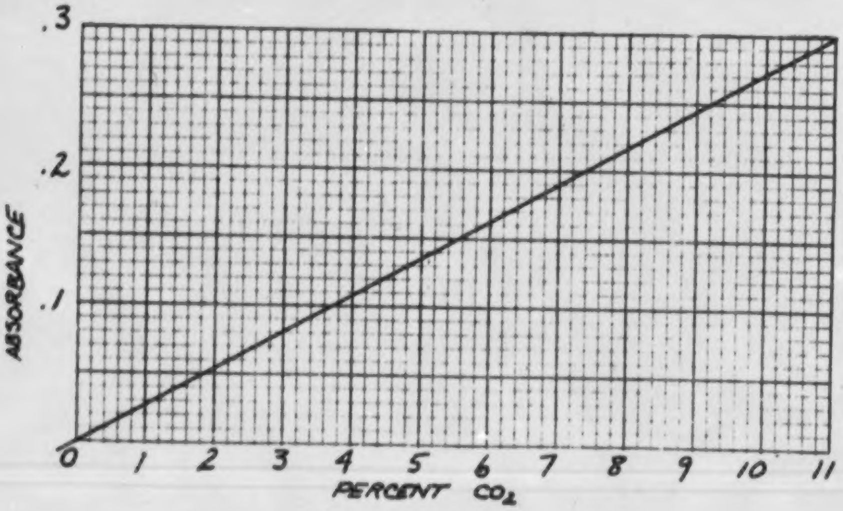
SAMPLE DATA SHEET:

| Tube # | Contents | % T | A | % $CO_2$ |
|---|---|---|---|---|
| 1 | blank | 100 | 0 | 0 |
| 2 | normal<br>deep breath | | | |
| 3 | breath held<br>for 20 sec | | | |

BIP # ____________

Test Well # ______

STANDARD GRAPH:



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# LABORATORY ACTIVITY 33:

# DIFFUSION

## INTRODUCTION:

The random movement of dissolved molecules or ions will cause them to spread throughout a solution until their concentration is the same everywhere. The net result, which is called diffusion, is the movement of particles from regions of high concentration to regions of low concentration.

In this activity, you will have an opportunity to observe the diffusion of colored particles throughout a container of water. You will also be able to test for the diffusion of particles through a membrane. The membrane is semipermeable, which means that the pores or holes in it are of a size that permits small particles to pass through it, but not large particles. Small and large particles will be placed on either the inside or the outside of the membrane and tests will be made to see which particles move through the membrane.

There are two main solutions you will be using. One is a protein solution made from the whites of eggs. The other is an amino acid solution made by dissolving glutamic acid in water. Proteins consist of chains of amino acids connected by chemical bonds. They are one of the substances which make up all of the cells in your body.

## MATERIALS:

- .1 M $CuSO_4$
- .01 M NaOH
- .1 M NaOH
- glutamic acid solution
- egg albumen solution
- beaker, 50 ml
- 4 beakers, 150-ml or 250-ml
- 2 pieces dialysis membrane, approximately 10 cm long
- 2 lengths of string, 15 to 20 cm long
- balance
- scoopula
- graduated cylinder, 100-ml
- pipet, 10-ml
- 2 funnels
- pH paper, 1-11 and color scale
- forceps
- 2 or more test tubes, 16 x 125 mm
- glass-marking pencil
- toothpick
- food coloring
- medicine dropper

## PROCEDURE:

### PART I: OBSERVING DIFFUSION

1. Fill a 50-ml beaker almost full of tap water. Place it where it will not be jostled or disturbed in any way.

2. Obtain a drop of food coloring on the end of a toothpick. Gently touch the food coloring to the surface of the water. (If you do not see a small colored region in the water, repeat this step.)

3. Observe the water for a minute or two. Record your observations and the time.

4. Leave the beaker for 30 minutes. Then record your final observations.

### PART II: PREPARING SOLUTIONS

1. Determine the volumes of .1 M $CuSO_4$ and water that are needed to make 10 ml of .01 M $CuSO_4$ solution. Record the volumes on your data sheet.

2. Check your results with the instructor. When you have approval, prepare 10 ml of .01 $CuSO_4$ solution in a labeled 50-ml beaker.

3. Add about 100 ml of tap water to a beaker. Use forceps to dip a small piece of pH paper into the water. Compare the color with the color chart. If the pH is lower than 8 or 9, add a few drops of .01 M NaOH, swirl the water and recheck the pH. Continue until the pH is 8 or 9. Label the beaker "Tap Water, pH = 8 (or 9)."

### PART III: SETTING UP THE EQUIPMENT

1. Cut two pieces of dialysis membrane, each 10 cm in length.

2. Although the dialysis membrane looks like a sheet of cellophane, it can be opened up into a tube (see Figure 1 on the next page). To open the tube, hold one of the cut ends of the membrane under running water for about 30 seconds, until you feel the sides come unstuck. Run water through the tube to open it completely. Tie a tight knot in one end, being careful not to puncture the membrane with a fingernail. Check for leaks as shown in Figure 1. Pour about 30 ml of water into the tube, hold the open end closed and squeeze gently on the water. Then empty the tube of water. If there is a leak, use a fresh membrane. Complete this step for both of your membranes.

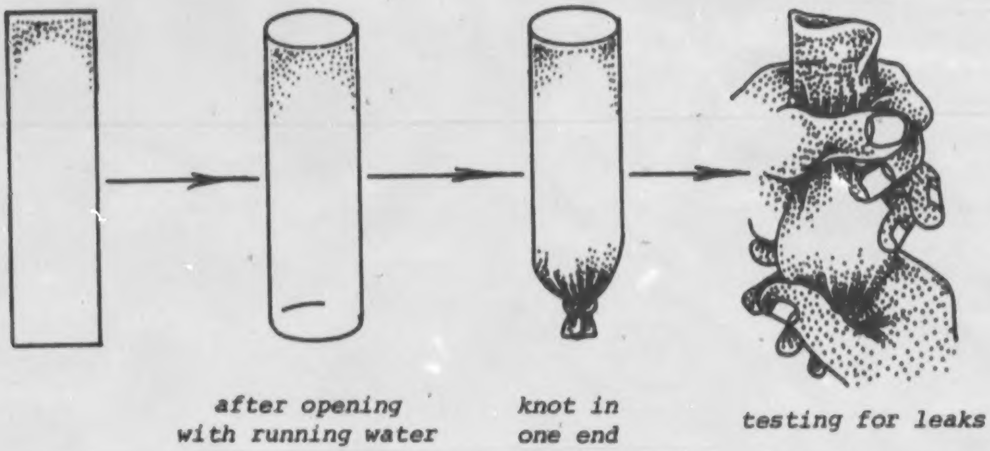


FIGURE 1: *Preparing and testing the dialysis membrane.*

3. Place a funnel into the open end of each membrane sac and tie the membrane securely to the funnel with a length of string (see Figure 2). Set each funnel and its membrane sac in a separate beaker, or suspend them from a ring-stand ring, as shown in Figure 2.

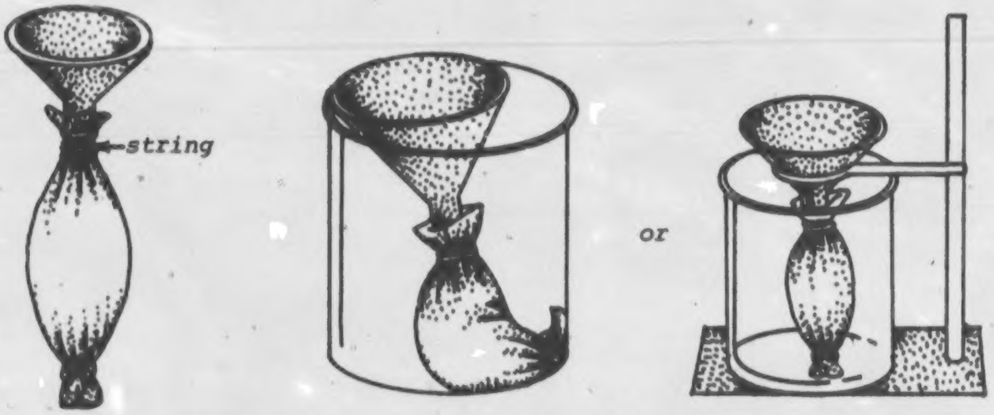


FIGURE 2: *Attaching and suspending the funnel.*

4. Into one funnel, pour about 5 ml of egg albumen solution and about 1 ml of .01 M $CuSO_4$ solution. Shake the funnel gently to mix the two solutions. Pour sufficient .1 M NaOH into the beaker to cover most of the membrane, but not the string holding it to the funnel. Swirl gently a few times. Label the beaker "protein" and set aside.

5. Into the second funnel, pour about 20 ml of glutamic acid solution. Into the beaker, pour the tap water that you adjusted to a pH of 8 or 9. Do not fill above the level of the string. Label the beaker "amino acid" and record the time.

## PART IV: CHECKING YOUR RESULTS

1. .01 M $CuSO_4$ and .1 M NaOH combine to make a reagent which tests for the presence of protein. The reagent is called biuret solution. To see what a positive test for protein looks like, you will need two test tubes. Label the tubes "1" and "2." In Tube #1, add 3 or 4 ml of .1 M NaOH. Then add about 10 drops of .01 M $CuSO_4$. Record the color on your data sheet. In Tube #2, add 3 or 4 ml of .1 M NaOH. Then add about 10 drops of .01 M $CuSO_4$ and about 10 drops of egg albumen solution. Record the color.

2. Watch the set-up with the egg albumen. If the color for the presence of protein appears either inside or outside the tube, you should record this fact. If the color for no protein appears, you should not that also.

3. Glutamic acid is an acid. When dissolved in water, it provides a pH lower than 7. If the glutamic acid molecules are small enough to move through the walls of the membrane, one might expect that the water in the beaker outside the tube would become lower in pH. About every 10 minutes, record the time and the pH of the water in the beaker. Gently swirl the beaker. Then use forceps to dip a small piece of pH paper in the water. Compare the color of the pH paper with the color chart. Continue to test the water for 30 to 40 minutes or until you run out of time. Record the result each time.

## DISCUSSION QUESTIONS:

1. In Part I, what caused the movement of the drop of food coloring?

2. In the set-ups with the egg albumen and glutamic acid, was there any evidence of a substance crossing the membrane? If so, what substance moved and in which direction?

3. What caused the substance(s) to move through the membrane?

4. Which molecules are too large to pass through the holes in the membrane--proteins or amino acids?

5. How do you think the concentration of glutamic acid inside the membrane would compare with the concentration outside, if the set-up were allowed to stand for a long time?

6. a. Recall what you have learned about acids and bases. Is it possible that a change in pH can occur without the amino acid itself moving through the membrane? Explain.

b. Did the procedure actually prove that glutamic acid passed through the membrane?

c. Discuss with other students how you might improve on the design of the experiment.

## SAMPLE DATA SHEET:

### PART I:

Initial time:

Observations:

Final time:

Observations:

### PART II:

To make 10 ml of .01 M $CuSO_4$:

Mix ______ ml of .1 M $CuSO_4$ + ______ ml of water.

### PART III:

Time (protein set-up):

Time (amino acid set-up):

### PART IV:

Tube #1: .01 M $CuSO_4$ + .1 M NaOH → ______________
color

Tube #2: .01 M $CuSO_4$ + .1 M NaOH + egg albumen (protein) → ______________
color

Protein Set-Up:

color inside:

color outside:

Amino Acid Set-Up:

Initial pH:

pH after 10 min:

pH after 20 min:

pH after 30 min:

## LABORATORY ACTIVITY 34:

## BLOOD, WATER AND OSMOSIS

### GENERAL INTRODUCTION:

If you have ever been a patient in a hospital, or visited one, perhaps you can recall the presence of bottles, each about a liter in volume, hanging above some of the beds. Some patients are fed liquid from these bottles directly into their bloodstream through a tube connecting the bottle to an arm vein (transfusion).

The fluid in the bottle is called normal saline. It has a particular concentration of salt. Whenever someone has much liquid added to their blood, it has to have that special salt concentration.

In this activity you will be presented with the puzzle of why the concentration of a fluid that is added to blood has to be so special. First, you will mix a drop of blood into some solutions. Each sample will contain a different salt concentration. Then you will observe the different samples with and without a microscope to see whether this provides any clues to the requirement for a particular salt concentration. Two principles are essential to the solution of the puzzle: diffusion, which was the subject of the last activity, and another phenomenon called osmosis, which may be demonstrated by the instructor.

This activity calls for two new tools. The first is called a lancet, which is used to obtain a drop of blood from the fingertip. This procedure is perfectly safe. The lancet will make a tiny cut just big enough for a drop to come out. It is already sterile so there is no need to worry about getting an infection from the lancet itself.

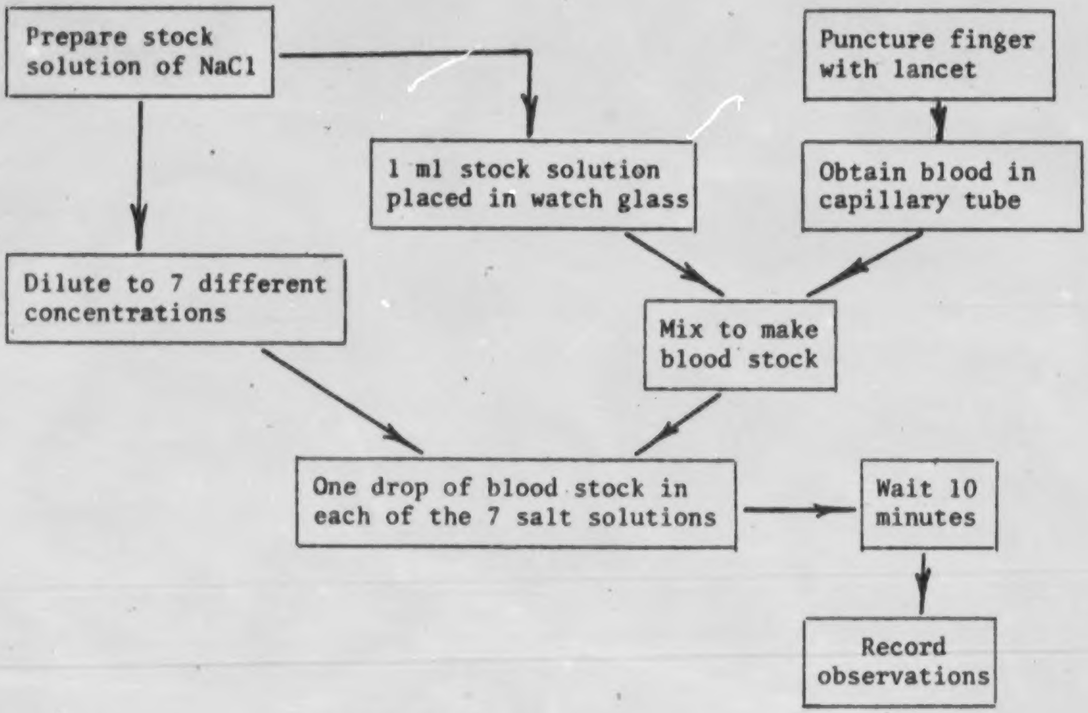
The other new tool is a capillary tube, which is a glass tube with a narrow diameter. When the end of such a tube is touched to a liquid, the liquid moves into the tube and may almost fill it up. This process, called capillary action, is the same process that causes water to fill a sponge that touches it.

## PART I: MIXING BLOOD WITH SOLUTIONS CONTAINING DIFFERENT SALT CONCENTRATIONS

### INTRODUCTION:

First, you will prepare seven test tubes, each with a different concentration of sodium chloride (NaCl). Next, you will obtain a drop of blood with the lancet, and use a capillary tube to transfer the blood to a watch glass containing a salt solution. A drop of this blood stock is then placed into each of the test tubes, with different concentrations, and observed.

The flow chart that follows shows the main steps of the procedure. Use it to get an overview of the activity before you begin.


Prepare stock solution of NaCl
Puncture finger with lancet
1 ml stock solution placed in watch glass
Obtain blood in capillary tube
Dilute to 7 different concentrations
Mix to make blood stock
One drop of blood stock in each of the 7 salt solutions
Wait 10 minutes
Record observations

FLOW CHART FOR PART I

MATERIALS:

test-tube rack
7 test tubes, 16 x 125 mm
1 g NaCl (sodium chloride)
2 beakers, 150-ml
glass rods
2 pipets, 1-ml
capillary tube, 1.2 x 75 mm
alcohol cotton wipe packet
lancet
glass-marking pencil
metric ruler
balance
graduated cylinder, 50-ml
watch glass
medicine dropper

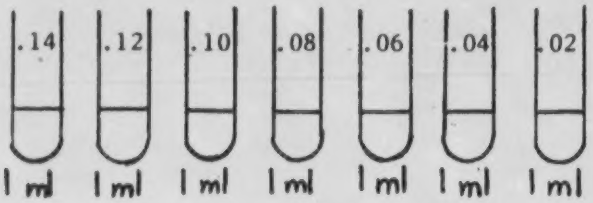
PROCEDURE:

1. Prepare 50 ml of a 0.14 M sodium chloride solution in a beaker. This is the stock salt solution.

2. The sodium chloride solution will be diluted to provide seven different concentrations. In the rack, set up seven test tubes in a row and label them .14, .12, .10, .08, .06, .04 and .02. These numbers stand for the molar concentration of the sodium chloride to be made in each.

To prepare the first test tube, which is to have a concentration of .14 M, one can simply add 1 ml of the stock salt solution. To prepare the .12 M tube one can put some of the stock salt solution into it and then add some water to dilute it slightly to make it .12 M. The smaller the salt concentration that is desired, the more tap water must be added to dilute it down.

The amounts used must be precisely measured. Figure 1 shows how much of the stock salt solution and how much water must be added to each test tube to make the sodium chloride concentrations you have written on the tubes.



| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| ROW A | amount of 0.14 molar sodium chloride to be added (ml) | 1.0 | .86 | .71 | .57 | .43 | .29 | .14 |
| ROW B | amount of water to be added (ml) | 0 | .14 | .29 | .43 | .57 | .71 | .86 |

FIGURE 1: *Volumes for seven concentrations of sodium chloride solution.*

Using a 1-ml pipet, first pipet the volume of stock salt solution indicated in Row A of Figure 1 into each of the seven test tubes. For example, the test tube labeled ".10" receives .71 ml of the stock salt solution. Then use a different 1-ml pipet to deliver the volume of water indicated in Row B into each test tube. This will result in the sodium chloride concentrations written on each tube.

3. Use the same pipet you used for the Row A volumes to pipet 1 ± 0.1 ml of the stock salt solution into a watch glass. Set it aside so that nothing will splash into it.

4. Now you're ready to get a drop of blood. Obtain an alcohol cotten wipe and a lancet, unopened. Also obtain a capillary tube. Review all the rest of this step before going on.

Choose someone to puncture your finger for you, unless you wish to do it yourself. (Note that most of the following instructions are for the person making the puncture.) First, examine--do not open--the package containing the sterile blood lancet. One end should read "open this end" or "peel to open." It is most important that you open that end, since it is the "handle" end. The opposite end is the blade which you should not touch, in order to keep it sterile.

The best finger to puncture is the one that gets the least use. This is usually the left ring finger (or the right ring finger for left-handed persons). Before the puncture is made, it is important to cleanse the skin. To do this, open the alcohol cotton wipe package. Then hold the finger with one hand and scrub its tip thoroughly with the alcohol wipe.

Open the blood lancet package. Make a quick, firm puncture with the lancet in the tip of the finger (see Figure 2), using a rapid in-and-out motion. If blood does not flow immediately, quickly repeat the puncture. Dispose of the lancet as directed by your instructor. Used lancets are contaminated and can cause infection if reused by someone else.

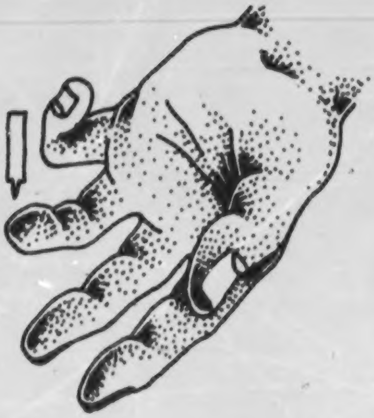


FIGURE 2: *Puncturing the fingertip.*

With the finger pointing downward, touch the end of the capillary tube to the drop of blood. The blood will quickly run into the tube. Allow enough blood to run in to form a column about 2 cm long. This amount can usually be obtained by placing your thumb at the joint nearest the nail and applying gentle pressure downward toward the wound. If you can't obtain a column 2 cm long, a column only 1.5 cm long will be enough. Don't let the blood remain in the capillary tube for too long, for it will clot.

5. Dip the end of the capillary tube into the salt solution in the watch glass. Very gently blow out the blood into the solution. Remove the capillary tube, and use a stirring rod to mix the blood and salt solution thoroughly but gently. The resulting mixture is the blood stock solution.

6. Now you are ready to place one drop of the blood stock into each of the seven prepared test tubes. Use a medicine dropper to release one drop of the blood stock directly into the liquid in the bottom of each test tube.

If you make a mistake, such as allowing more than one drop to fall into a test tube, simply proceed to the remaining test tubes. Then return to the tube with the error and empty it. Repeat Steps 2 and 6 for that tube.

7. Jiggle each tube a bit to distribute the blood stock uniformly throughout. Wait 10 minutes.

8. Most of the tubes will probably be cloudy, but a few should appear clear. To observe this, hold the tubes up to the light. Compare each tube with the one labeled .14. Record your results.

SAMPLE DATA TABLE:

| Molar concentration of blood solution | Appearance: clear or cloudy | Further comments |
|---|---|---|
| .14 | | |
| .12 | | |

## PART II: OBSERVING BLOOD CELLS WITH THE MICROSCOPE

INTRODUCTION:

In Part I you probably observed that blood reacted differently to different concentrations of salt solution. This observation should provide a clue to the original puzzle--why a transfusion uses a salt solution of a particular concentration. But what is actually happening to the blood in the different solutions? In Part II you will study blood at the cellular level to answer this question.

MATERIALS:

| | |
|---|---|
| microscope | up to 7 cover slips |
| up to 7 glass slides | pipet, 1-ml |

PROCEDURE:

1. Pipet a drop of the blood solution from the test tube labeled .14 onto a slide. Then carefully place a cover slip on the drop in order to spread it out. (You may want to review Laboratory Activity 10 on the use of the cover slip and microscope.)

2. Examine the blood solution using the low-power objective. Use the diaphragm to cut down the light; you won't be able to see blood cells if the field is too bright. Focus until you find red blood cells. These cells will look something like small faint circles. The other blood cells are much less common and are not likely to be seen.

3. In a similar manner, prepare slides of another one of the cloudy tubes. Then prepare a slide of one or more of the clear tubes. Look for blood cells. Do you see cells in all of the cloudy tubes? Do you see cells in all of the clear tubes? Record your observations.

SAMPLE DATA TABLE:

| Molar concentration of blood solution | Red blood cells--present or absent | Further comments |
|---|---|---|
| .14 | | |
| | | |

DISCUSSION QUESTIONS:

1. Which salt concentrations showed a clear solution with blood? In looking across the row of tubes, does one tube suddenly appear clear; or is there a gradual change from cloudiness to clarity?

2. What did you observe with the microscope? How did these results compare with the appearance of the solutions in the seven test tubes?

3. What do the microscopic observations suggest happened to the red blood cells?

4. What explanation can be offered for what happened to the red blood cells? (Hint: The principle of osmosis is involved. The red blood cells absorbed water in varying amounts, depending on the salt concentration around them. Each red blood cell has a cell membrane completely enclosing the water inside it. This membrane is rather fragile.)

5. Under which circumstances did the red blood cells absorb more water, when they were placed in a high salt concentration or a low salt concentration?

6. Why must liquid that is added to a person's bloodstream have a special salt concentration? What concentration of salt would you be sure not to add to someone's blood?

## LABORATORY ACTIVITY 35:

## THE ELECTROLYSIS OF WATER

### GENERAL INTRODUCTION:

Water is the most common of all biological substances. It comprises approximately three-fourths of the total mass of our bodies. It is a very special substance, since it is the essential and fundamental matrix for all known life processes. In other words, most chemical reactions which take place in the body occur in some type of water environment.

As is commonly known, the formula for water is $H_2O$ which means that it is composed of hydrogen and oxygen atoms. However, oxygen and hydrogen exist in their elemental states as gaseous molecules ($H_2$ and $O_2$). These substances bear little resemblance to water. The unique properties of water are derived, in part, from the type of molecular bonding which occurs between the atoms of hydrogen and oxygen.

In this activity, these molecular bonds are broken by passing an electrical current through water. (Actually, a dilute ionic solution is used. As demonstrated in Laboratory Activity 24, pure water is a very poor conductor of electricity.) This process is known as electrolysis; the products of the reaction are hydrogen and oxygen gas. In addition, an alternate method for producing hydrogen gas is provided. Using various tests, the properties of hydrogen in its elemental state may be investigated.

## PART I: THE ELECTROLYSIS OF WATER

### INTRODUCTION:

In this part of the activity, the electrolysis apparatus is set up and operated. Samples of oxygen and hydrogen gas are collected, which are tested in Part III.

### MATERIALS:

beaker, 250-ml
275 ml electrolysis solution
2 electrodes
BIP
ring stand
ring-stand clamp
4 to 6 test tubes, 16 x 125 mm, with stoppers
test-tube rack
clock with second hand
glass-marking pencil

**PROCEDURE:**

1. Obtain approximately 100 ml of electrolysis solution in a 250-ml beaker.

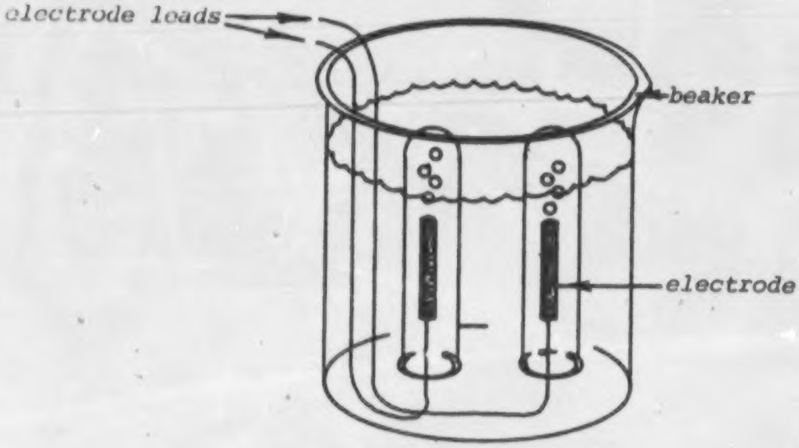
2. Place the test tubes in the rack and fill them to the brim with this solution. These tubes are used in Step 7 to collect samples of gas resulting from the electrolysis.

3. Refill the beaker until it contains approximately 175 ml of the electrolysis solution.

4. Place the electrodes in the beaker.

5. Connect the leads of the electrodes to pins S and W on the programming panel of the BIP. In addition, connect pin T to V.

6. Plug in the BIP, and observe the electrodes. If the connections are correct, gas should be bubbling from the electrodes.



*Collecting the gases.*

7. Collect 2 to 3 samples of the gas coming from each electrode as shown in the diagram above. This may be done as follows. (See Step 9 before completing this step. Also, it may be desirable for a partner to begin Part II, particularly if the gas is being produced slowly by the electrolysis apparatus.)

   a. Place a finger over the opening of one of the test tubes filled with electrolysis solution.

   b. Invert the test tube and position the opening over one of the bubbling electrodes.

c. Remove your finger without allowing any solution to flow out of the tube. Allow the gas to begin filling the tube.

d. Rinse your fingers in water to remove the electrolysis solution.

e. When the tube is filled with gas, carefully insert a stopper in the opening so that no gas escapes.

f. Label the test tube so that you know which electrode the gas came from. (Was the electrode connected to pin S or pin W?

g. Set the tubes aside for testing in Part III.

8. Record the lengths of time required to fill the test tubes for each electrode. Does one of the electrodes produce gas more rapidly than the other? How can this be accounted for?

## PART II: PREPARATION OF HYDROGEN GAS

### INTRODUCTION:

Hydrogen gas may also be prepared by reacting certain metals with dilute acid solutions. In this part of the activity, zinc metal (Zn) reacts with hydrochloric acid to produce zinc chloride ($ZnCl_2$) and hydrogen gas. The gas samples are tested in Part III.

### MATERIALS:

zinc metal
1 M hydrochloric acid solution
gas-generating appartus
(see p. 35)

### PROCEDURE:

1. Assemble the gas-generating apparatus shown on page 35. (The gas burner is not needed.)

2. Place 4 or 5 pieces of zinc metal in the test tube. The zinc may be safely handled with your fingers. Add about 25 ml of 1 M HCl and stopper the tube. Wait about 30 seconds.

3. Collect the gas just as you did in Laboratory Activities 8 and 9.

4. Collect several tubes of hydrogen gas. Label the tubes.

## PART III: TESTING HYDROGEN GAS

### INTRODUCTION:

In this part of the activity, the collected samples are subjected to certain tests. These tests demonstrate the properties of the gases produced by the electrolysis of water.

### MATERIALS:

tubes containing gas samples
wooden splint
matches

### PROCEDURE:

1. Take one of the tubes of hydrogen gas produced in Part II and hold it so that the opening is pointed down. (Can you think of a reason why this is necessary?)

2. Unstopper the tube and hold a burning splint near the opening. A distinct "pop" should be heard which is characteristic of hydrogen burning in a test tube.

3. Repeat this test on a sample of gas from each electrode of the electrolysis apparatus. From your results, can you tell which electrode produced the hydrogen gas? Does the gas from the other electrode behave like oxygen gas?

4. Test a tube of hydrogen gas using a glowing splint (one that is not burning). Does the hydrogen gas ignite the splint as oxygen gas does?

### DISCUSSION QUESTIONS:

1. Compare the behavior of hydrogen and oxygen when subjected to the glowing-splint and burning-splint tests.

2. Write a balanced equation for the electrolysis of water.

3. Write a balanced equation for the burning of hydrogen in air.

4. Write a balanced equation for the generation of hydrogen from hydrochloric acid and zinc.

5. Which of the two pins, S and W, on the BIP programming panel was connected to the anode? Which to the cathode? Explain your reasoning.

# LABORATORY ACTIVITY 36:

## CHANGING THE COLOR OF BLOOD

### GENERAL INTRODUCTION:

Everyone knows that blood is bright red--at least when coming from a cut. However, people with pale skin can see that certain blood vessels are more blue than red, such as the vessels on the inner surface of the wrist. The blood in these vessels has given up its oxygen to body tissue and has lost its red color. In some ways, the hemoglobin in the blood is similar to an indicator--when it is oxygenated it is bright red, and when it loses its oxygen it becomes bluish-purple.

The purpose of this activity is to measure how the blood changes color in response to circulation impairment or breath-holding. The changing color of blood may be detected using a device called an oximeter. Just as a colorimeter measures the amount of color in a solution, an oximeter measures the amount of oxygen present in the blood.

The oximeter provides information which can be used to determine how rapidly the blood circulates in the body. Measurement of circulation time is often done on patients with heart disease or circulation disorders. Oximeters can also be used to detect the onset of respiratory problems. A marked reduction of blood oxygenation almost always indicates a severe problem needing immediate medical attention.

## PART I: TESTING THE OXIMETER

### INTRODUCTION:

The principal electrical components of the oximeter are very similar to those of the colorimeter. The oximeter consists of a light source and a photocell which are attached to the BIP. The amount of light striking the photocell induces a flow of electricity (electric current) which can be measured using the BIP. This part of the activity demonstrates some of the principles of photocell operation in the colorimeter and the oximeter.

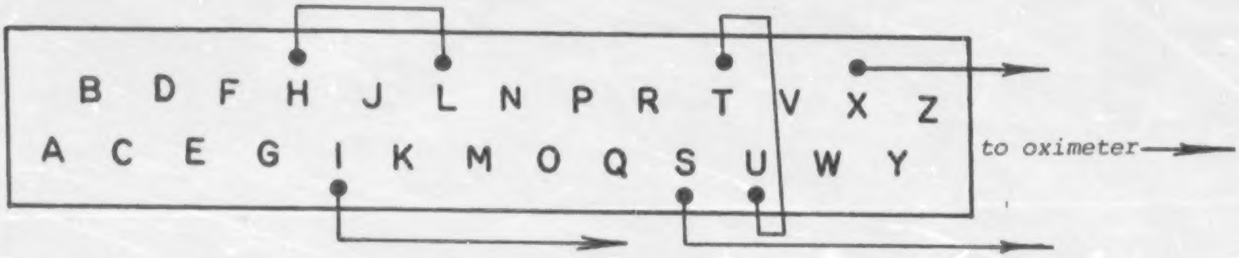
### MATERIALS:

oximeter
BIP

PROCEDURE:

1. Examine the oximeter. It consists of a light source (identical to the diode lights of the BIP) and a disc-shaped photocell. The photocell should be handled carefully to avoid scratching the sensitive surface.

2. Program the BIP as shown in the diagram. The three wires attached to the oximeter are labelled with letters that correspond to the pins on the programming panel of the BIP. Note that the wiring is identical to the programming of the colorimeter.



3. Plug in the BIP. The light in the oximeter should be "on" and about as bright as one of the BIP diode lights.

4. Cover the photocell so that it is completely shielded from light. For example, the photocell surface may be covered and pressed firmly against the palm of the hand or other part of the body.

5. Turn the mA dial until the diode lights indicate a null condition. If no light reaches the photocell, the reading should be near zero.

6. Place the fleshy part of the thumb over the photocell and readjust the mA dial. The reading should increase. The thumb is translucent enough so that some of the light is able to pass through, although not very much.

7. Vary the amount of light striking the photocell by shading it, or exposing it directly to room light. How does the amount of light reaching the photocell affect the BIP reading?

8. In very bright light, it may not be possible to adjust the mA dial to a "null" condition. This is because the photocell is producing more current than the BIP can measure. To compensate for this, remove the wire at L on the programming panel and insert it at M, N, O or P. These connections allow the BIP to measure increasingly greater amounts of light.

9. Insert a screwdriver into the slotted control of the voltage regulator, which is just below and to the left of the mA dial. Turn the screwdriver in either direction, and watch the diode light of the oximeter. If observed closely, it can be seen that the light dims and brightens, depending upon which way the control is turned. (In the colorimeter, this control adjusts the brightness of a light source inside the test well. Since this changes the amount of light hitting the photocell, this control is used to standardize the colorimeter as in Laboratory Activity 31, Step 14.

10. Turn the screwdriver full clockwise (the brightest condition of the diode light). Remove the screwdriver from the control. This control will not need further adjustment during the activity.

## PART II: USING THE OXIMETER

### INTRODUCTION:

Blood that has lost its oxygen to the body tissues is bluish-purple, and can absorb more light than red, oxygenated blood. In this part of the activity, the oximeter is used to detect these changes. The oximeter is positioned in such a way that light diode shines through the thumb. The amount of light transmitted is picked up by the photocell and measured with the BIP.

### MATERIALS:

BIP
oximeter
ACE bandage

### PROCEDURE:

1. The wire shown at L in the wiring diagram should be removed and inserted at K. This makes the measuring system inside the BIP more sensitive. This is necessary since the thumb transmits only very small amounts of light.

2. Be sure that your thumb is clean. Place the oximeter assembly over your thumb as shown in the diagram on the following page. The photocell should rest at the base of the thumb-nail. The diode light is pressed against the fleshy side of the thumb, directly opposite the photocell.

3. Wrap an ACE bandage around the assembly so that the oximeter is firmly held in place. (Don't wrap the bandage too tightly since this can reduce circulation to the thumb.) Wrap the bandage so that no light can reach the oximeter.

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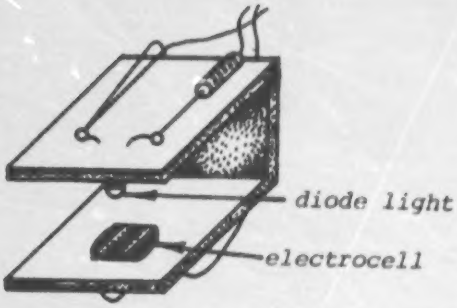
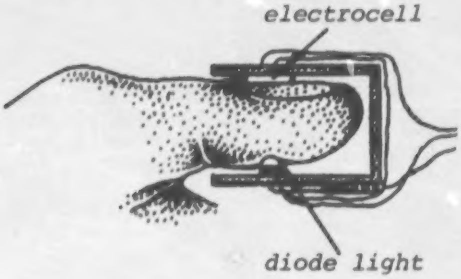




4. Turn the mA dial to the null point. It should be somewhere between 10 and 80. If the reading is less than 10, reposition the oximeter so that the null point is higher. The BIP oximeter is now measuring the light from the light diode that is passing through your thumb and reaching the photocell on the other side. To prevent uncertain results, keep your thumb still while making measurements.

5. Hold your breath for 30 seconds. Have another student make a reading within a few seconds after you finish holding your breath. Repeat the procedure, holding your breath for longer periods of time, up to the maximum amount of time that you can reasonably hold your breath. Record your results.

6. After you have become familiar with the oximeter, see whether you can solve the following problem. Devise a way of using the oximeter to measure the time it takes blood to get from your lungs to your thumb. Record your method in your data sheet, and then try it out. Enter your results on your data sheet.

## LABORATORY ACTIVITY 38:

## MEASURING ONE-SECOND FORCED EXPIRATORY VOLUME ($FEV_1$)

### INTRODUCTION:

Healthy individuals have the ability to exhale almost all of the air from their lungs very quickly. This ability allows the body to respond to extreme oxygen demands whenever necessary (such as during vigorous exercise).

Persons with impaired lung function, such as occurs in chronic bronchitis, asthma and emphysema, are not able to exhale quickly. The most important test of pulmonary function is the one-second forced expiratory volume ($FEV_1$). This is the volume of air that a person can expire in one second, following maximum inspiration.

$FEV_1$ is usually expressed as a percentage of the total forced expiratory volume, which you measured earlier. The total FEV is often written as $FEV_T$ to distinguish it clearly from $FEV_1$.

$$\frac{FEV_1}{FEV_T} \times 100 = \text{percentage of vital capacity that is expired in one second}$$

In healthy young persons, this value is typically around 90%. The percentage drops off gradually with age, reaching around 75% in healthy elderly persons.

In this activity, you will measure $FEV_1$ and $FEV_T$ and calculate the percentage just described. The procedure requires close coordination between you and a second person. Both of you should read the entire procedure carefully before beginning the activity.

### MATERIALS:

source of timed, one-second signals
spirometer
plastic wrap or other suitable mouthpiece covering

### PROCEDURE:

1. Obtaining a measure of $FEV_1$ is different in two respects from the procedure used to measure $FEV_T$ in Laboratory Activity 18. First, the subject should expel the air from his or her lungs as rapidly as possible. Second, the subject and tester must have a pre-arranged starting time for the exhalation; and the tester must obtain a spirometer reading while the inner container is still in motion.

2. The instructor will provide a source of signals that occur once each second. The subject and tester should agree on a scheme that permits both of them to know on which signal the expiration will begin. For example, "One, two, three, four, blow!" called off by the tester in time with the one-second signals. During the first three counts, the subject fills his or her lungs as completely as possible. On the count of four, the subject's breath is being held; and on the fifth signal, the expiration begins.

3. The tester must be in a position to estimate the position on the spirometer scale just as the sixth signal occurs. This volume should be recorded as $FEV_1$. The tester should also estimate the imprecision involved in taking a reading while the spirometer is in motion. Might the reading be off by 0.1 liter in either direction? 0.2 liter? Record $FEV_1$ as a range of imprecision--for example, 5.4 ± 0.2 liters.

4. The subject should continue to exhale until the lungs have been emptied as completely as possible. Then either the subject or the tester should record the final reading as $FEV_T$.

5. Repeat the procedure twice more so that there are three pairs of values for the subject.

6. Use the range of imprecision for $FEV_1$ to determine a maximum and minimum value for $FEV_1$ expressed as a percentage of $FEV_T$. Then express the percentage as a range of imprecision. For example, assume that the first trial gave $FEV_1 = 5.4 \pm 0.2$ liters and $FEV_T = 5.9$ liters. We ignore the imprecision of the $FEV_T$ measurement, because it is small in relation to the imprecision of the $FEV_1$ measurement.

$$\text{maximum percentage} = \frac{5.4 + 0.2}{5.9} \times 100$$

$$= \frac{5.6}{5.9} \times 100$$

$$\simeq 94.9\%$$

$$\text{minimum percentage} = \frac{5.4 - 0.2}{5.9} \times 100$$

$$= \frac{5.2}{5.9} \times 100$$

$$\simeq 88.1\%$$

The midpoint of the percentage range is

$$\frac{94.9 + 88.1}{2} = \frac{183.0}{2} = 91.5\%$$

The imprecision is

$$\frac{94.9 - 88.1}{2} = \frac{6.8}{2} = 3.4\%$$

So the range of imprecision for the percentage is 91.5 ± 3.4%

DISCUSSION QUESTIONS:

1. Why do you think $FEV_1$ expressed as a percentage of $FEV_T$ is a better measure of lung function that $FEV_1$ by itself?

2. Compare your results with those of other class members. If there are smokers in your class, compare their results with those of non-smokers. Is there a noticeable difference?

## LABORATORY ACTIVITY 39:

## MEASURING $NO_2$ CONCENTRATIONS IN AUTOMOBILE EXHAUST

### INTRODUCTION:

Nitrogen oxides are among the important air pollutants present in city environments. Whenever burning occurs, especially under conditions of high temperature and pressure, a small amount of atmospheric nitrogen burns as well. This is represented in the following equations.

$N_2 + O_2 \longrightarrow 2\ NO$ nitric oxide gas

$N_2 + 2\ O_2 \longrightarrow 2\ NO_2$ nitrogen dioxide gas

Relatively harmless in itself, nitric oxide in the presence of sunlight reacts with oxygen to form nitrogen dioxide.

$$2\ NO + O_2 \longrightarrow 2\ NO_2$$

$NO_2$ is a brown colored gas. It is seen as a brownish haze in polluted areas. About 50% of the $NO_2$ present in cities is produced by automobiles.

The major effects of $NO_2$ pollution are reduction of visibility, damage to vegetation and eye irritation. In addition, $NO_2$ can react with other gases in the presence of sunshine to form photochemical smog.

Nitrogen dioxide concentrations in the atmosphere are usually measured in units of parts per million (abbreviated ppm). This refers to the number of parts by volume of $NO_2$ in a million parts of air. Nitrogen dioxide concentrations usually range from 0 to .20 ppm except in extreme cases. The following table lists typical air quality designations used with different levels of $NO_2$.

| | |
|---|---|
| clean air | .00 to .10 ppm |
| light air pollution | .11 to .14 |
| significant air pollution | .15 to .20 |
| heavy air pollution | .21 to .30 |
| severe air pollution | > .30 |
| emergency | > 1.60 |

The purpose of this activity is to measure the concentration of nitrogen dioxide in exhaust gases. A measured volume of exhaust gas is exposed to an absorbing reagent that reacts with $NO_2$ to produce a red color. The amount of color is directly related to the concentration of $NO_2$ and is measured using the BIP colorimeter.

## PART I: PREPARING A STANDARD GRAPH

In order to measure unknown $NO_2$ concentrations in exhaust gases, it is necessary to prepare known reference solutions. Sodium nitrite ($NaNO_2$) is used since it reacts with the absorbing reagent in the same manner as nitrogen dioxide gas. The absorbing reagent reacts with the $NO_2$ group in either the ionized $NO_2^-$ (nitrite ion) form or in the molecular (gaseous) form.

### MATERIALS:

25 ml absorbing reagent
10 ml sodium nitrite solution
5 test tubes, 16 x 125 mm, with stoppers
pipet, 10-ml
pipet, 1-ml
2 beakers, 50-ml
glass-marking pencil
BIP colorimeter
Absorbance-Transmittance Table
test-tube rack

### PROCEDURE:

1. Program the BIP for use as a colorimeter. Allow it to warm up for at least 10 minutes.

2. Obtain about 10 ml of the sodium nitrite solution in a 50-ml beaker.

3. Obtain about 25 ml of absorbing reagent in a 50-ml beaker. Label the beakers so that they may be distinguished.

4. Prepare the dilutions that are listed in the following table. For example, 0.2 ml sodium nitrite solution and 4.8 ml of absorbing reagent should be mixed in Tube #1. Be sure to use the 10-ml pipet only for the absorbing reagent and the 1-ml pipet only for the $NaNO_2$ solution.

| Tube | ml Absorbing Reagent | ml $NaNO_2$ Solution | %T | A | PPM |
|---|---|---|---|---|---|
| 0 | 5.0 | 0.0 | 100 | .00 | 0.0 |
| 1 | 4.8 | 0.2 | | | 0.5 |
| 2 | 4.6 | 0.4 | | | 1.0 |
| 3 | 4.4 | 0.6 | | | 1.5 |
| 4 | 4.2 | 0.8 | | | 2.0 |

Note that 0.4 ml of the $NaNO_2$ solution gives an absorbance equal to that of $10^{-6}$ liter of $NO_2$ gas.

5. Stopper the tubes and invert them once or twice to mix the contents.

6. Tube #0, which contains only the absorbing reagent, is used as the blank. With this test tube in the colorimeter test well, the mA dial is set at 100, and the variable voltage control adjusted until the light diodes null. (The variable voltage control is the slotted control just below and to the left of the mA dial.)

7. Measure and record the % transmittance readings of the remaining four tubes.

8. Convert the % transmittance readings to absorbance values using the Absorbance-Transmittance Table.

9. On a sheet of graph paper, plot absorbance as a function of concentration in ppm (concentration on the horizontal axis). ~~Then~~ draw a "best" straight line that passes through or near the origin (0,0) and as close as possible to the other data points. Extend the line to include absorbances up to .60.

## PART II: MEASURING $NO_2$ CONCENTRATIONS IN COLLECTED GAS SAMPLES

In the second part of this activity, exhaust samples are collected in plastic bags. The samples are transferred into bottles that contain a small amount of absorbing reagent. The bottle is capped and shaken for several minutes. This permits the $NO_2$ in the exhaust gas to react with the absorbing reagent, producing the red color.

The $NO_2$ concentration is measured by comparing the absorbance of the red solution with the standard graph made in Part I. Samples that produce more color than can be measured accurately (absorbances greater than 0.6) must be diluted. It is important to keep the dilution factors in mind while making calculations.

The flow chart on the next page provides an outline of the procedure.

## MATERIALS:

The materials list that follows indicates the equipment necessary to perform one test. For additional tests, certain materials are re-used, such as the bottle, plastic bag and colorimeter.

20 ml absorbing reagent
2 test tubes, 16 x 125-mm with stoppers

2 pipets, 10-ml

beaker, 50-ml

BIP colorimeter

test-tube rack

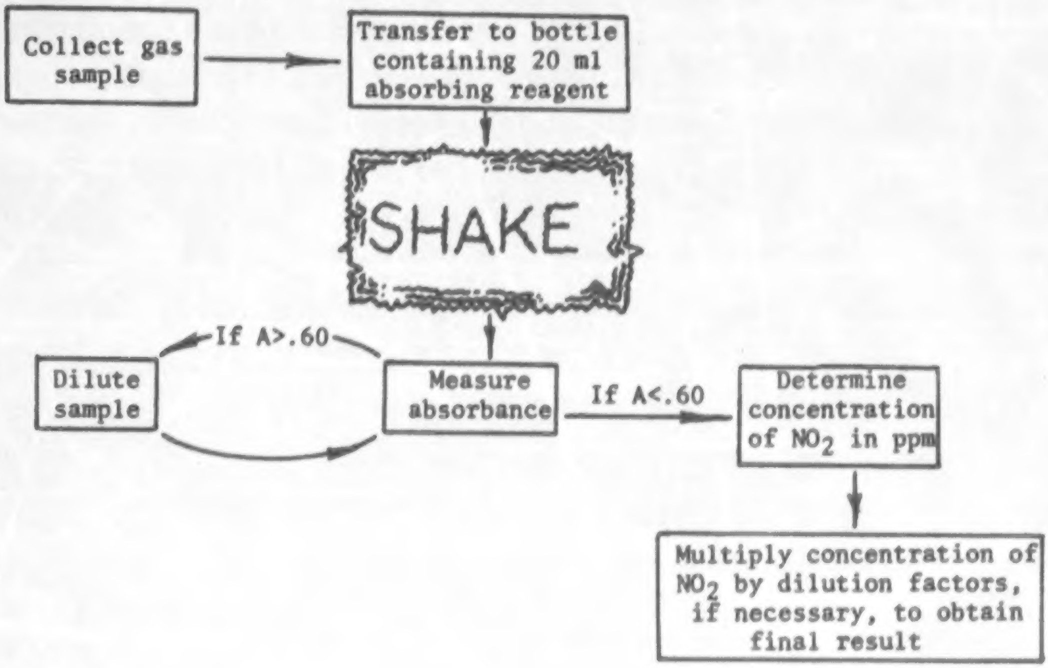
one-gallon bottle with cap

plastic bag with twist tie

short length (10 to 15 cm) tubing, glass, rubber or plastic, 12 to 20 mm in diameter

graduated cylinder, 100-ml

rubber band


Collect gas sample
Transfer to bottle containing 20 ml absorbing reagent
SHAKE
If A>.60
Dilute sample
Measure absorbance
If A<.60
Determine concentration of $NO_2$ in ppm
Multiply concentration of $NO_2$ by dilution factors, if necessary, to obtain final result

PROCEDURE:

1. Collect a sample of exhaust gas in a plastic bag by placing the opening of the empty bag over the exhaust pipe of the vehicle being tested. When the bag is filled, quickly close the opening and secure it with the twist tie. (Caution: the exhaust pipe of a car that has been running for more than a few minutes will be too hot to touch. It is only possible to take samples in the manner described if the engine is moderately cool.)

2. Attach a piece of tubing to the mouth of the bag with a rubber band without undoing the tie. (See Figure 1.)

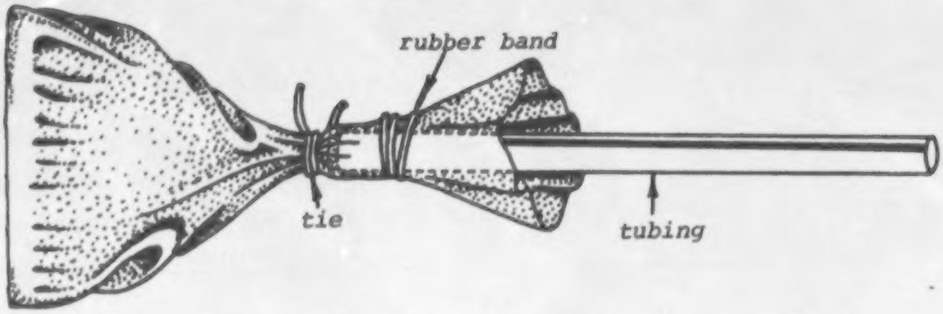


FIGURE 1: *Attaching tubing to the plastic bag.*

3. With a 100-ml graduate, measure out 20 ± 0.5 ml of absorbing reagent and place it in a clean one-gallon bottle. (If several tests are done at a time, it is difficult to dry the bottle completely before each use. For the purposes of this procedure, it is acceptable to rinse the bottle briefly with tap water after each test. The bottle should be inverted over a paper towel for a few minutes so that most of the water drains out.)

4. Insert the tubing that is connected to the plastic bag into the mouth of the bottle. Undo the tie and squeeze the contents of the bag into the bottle. (Since the bag contains a larger volume than the bottle, almost all of the air initially present in the bottle is blown out.) See Figure 2 below.

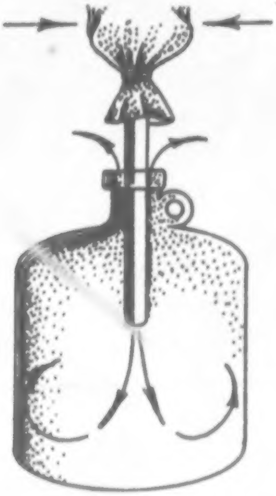


FIGURE 2: *Emptying the exhaust gases into the bottle.*

5. Remove the tubing and quickly cap the bottle. The bottle must be shaken for 2 to 4 minutes so that all of the $NO_2$ present is exposed to the absorbing reagent. The more the mixture is shaken, the better the accuracy of the results.

6. Pour the solution out of the bottle and into a small beaker.

7. Place about 10 ml of the unknown in a test tube, measure the % transmittance and determine the absorbance. If the absorbance is greater than 0.60, go on to Step 8. If the absorbance is less than 0.60, skip Step 8 and go on to Step 9.

8. If the absorbance is greater than .60, the solution should be diluted. First try diluting the sample by a factor of 10 by adding 9 ml of water to 1 ml of the unknown in another test tube. Remember to keep track of any dilutions that are made. It may be necessary to dilute the sample further in order to bring it into the proper range.

9. If the absorbance falls within an acceptable range, use the graph you made in Part I to find the concentration in ppm of the unknown. Multiply the result by any dilution factors from Step 8 to obtain the concentration of $NO_2$ in the exhaust gas. For example, if the unknown was diluted by a factor of 10, the result obtained from the graph must be multiplied by 10.

## DISCUSSION QUESTIONS:

1. Compare your results with those obtained by other members of the class. Does the age of the car seem to make a difference? What about engine size? Do cars that sound as if they are running poorly produce more $NO_2$?

2. Why was it necessary to shake the bottle containing the exhaust gases for such a long time?

3. Why is a dilution necessary in the case of an unknown that gives an absorbance greater than .60? Why was the line extended to .60, but not beyond?

## LABORATORY ACTIVITY 40:

## THE EFFECT OF TEMPERATURE ON SMOG CONDITIONS

### INTRODUCTION:

Under normal conditions, atmospheric temperature decreases with altitude (approximately 3 °C decrease per 1,000 feet of elevation). The rise of warmer, less dense air results in the vertical mixing of air, which is an important factor in the dispersion of air pollutants.

When the air at or near ground level is cooler than the atmosphere above, thermal inversions can result. This can lead to the occurrence of smog conditions, since little or no vertical mixing of air takes place. The objective of this activity is to investigate the effects of temperature on vertical air movement.

### MATERIALS:

acetate sheet (clear),
  approximately 60 x 150 cm
rubber bulb with one-way valve
Tygon tubing, 40-cm length
Tygon tubing, 5-cm length
pan or other water container
ice
3 thermometers, 0-100 °C
2 cigarettes, filter-tip
string (approximately 100 cm)
2 paperclips
2 beakers, 250-ml
length of glass tubing
masking tape or large rubber bands
gas burner
matches
small can (for ashtray)

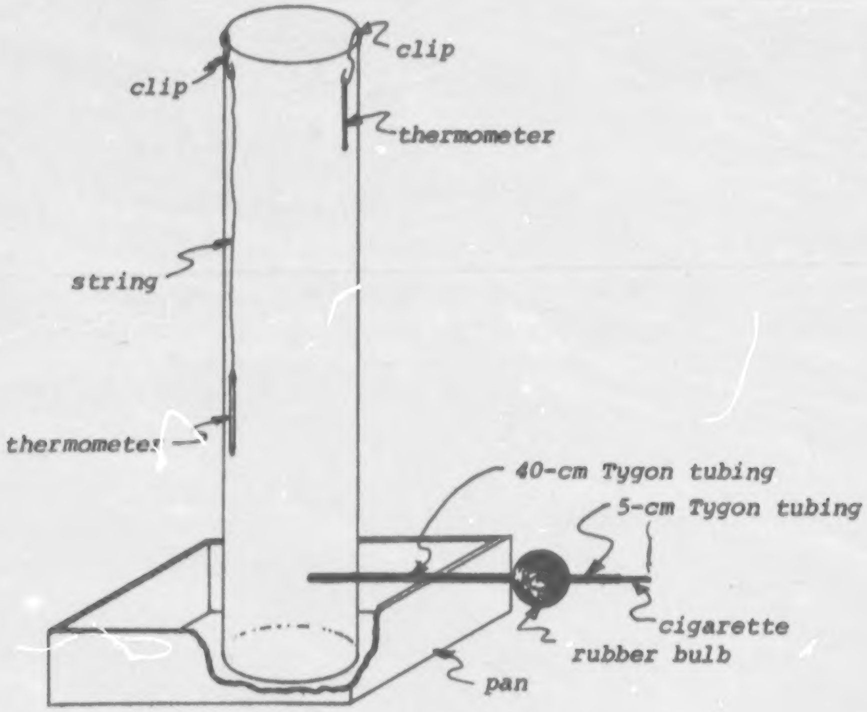
### PROCEDURE:

1. Assemble the equipment as shown in the figure on the next page. Start by rolling the acetate sheet into a cylinder 120 cm high and 60 cm in circumference. The two edges may be held together by sealing them with masking tape or by wrapping large rubber bands around the cylinder at several points along its length.

2. A hole large enough to accept the tubing may be melted in the side of the acetate cylinder by using the end of a piece of glass tubing that has been heated in a burner flame. Note that the hole *must* by high enough so that it will be above the water level when the container is *completely filled*.

3. The rubber bulb pumps air in only one direction. Experiment with it to determine which end sucks air in and which end blows it out.

4. Attach the "blowing" end of the bulb to the long piece of Tygon tubing. Attach the "sucking" end of the bulb to the short piece of tubing.



5. Insert the long piece of tubing into the hole in the wall of the acetate cylinder.

6. Position the cylinder in the water container so that the rubber bulb is outside the container.

7. Fill the container about half full of 45 °C tap water. (This can be done with a mixture of hot and cold water.) Maintain a water temperature of approximately 45 °C throughout this phase of the activity by adding hot water as needed.

8. Open a paperclip into an "s" shape. Use the paperclip and string to hang a thermometer inside the cylinder so that the bulb is about two-thirds of the way down the length of the cylinder.

9. Use a second paperclip, without string, to hang a second thermometer as high in the cylinder as possible.

10. Position both thermometers so that they can be read through the cylinder wall.

11. After the thermometers have adjusted to the new temperatures, record the locations of both thermometer bulbs and the temperatures at both positions.

12. Record the temperature of the water in the container.

13. Insert the filter tip of a lighted cigarette into the open end of the short length of tubing.

14. Alternately squeeze and release the rubber bulb so that smoke is pumped into the air-movement column. Observe the movement of the smoke. What does the behavior of the smoke indicate about the movement of air within the cylinder? (1) Why does the air move in this manner? (2)

15. Remove the thermometers and acetate cylinder. Remove the smoke from the column by passing it gently through the air of the room. Discard the water in the pan.

16. Re-assemble the apparatus. Fill the container about half full of 8 °C tap water (by mixing ice and cold tap water). Maintain a water temperature of approximately 8 °C throughout this phase of the activity by adding ice as needed.

17. Repeat the procedure at the new temperatures. Describe the behavior of the smoke in this case. (3) Explain why the smoke behaves in this way. (4)

DISCUSSION QUESTIONS:

5. Discuss whether this activity reproduced conditions found in the natural environment. Include in your answer some of the atmospheric conditions that can bring about air inversions.

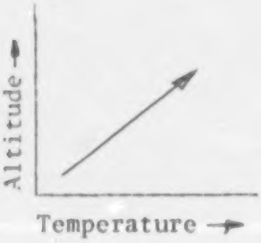
6. Would you expect the air movement to be affected by the diameter or the length of the column? Explain and/or test your answer.

7. A mass of air at ground level has a volume of 40 liters and a pressure of P. Assume that the air mass is moved to an altitude such that the new pressure is $\frac{1}{2}$ P. If the temperature remains constant, what is the new volume of the air mass?

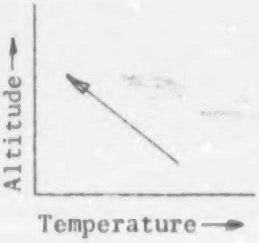
8. The mass of the air in the preceding question is 1,200 grams. What is its density at ground level? What is its density at the new altitude?

9. Which of the following temperature profiles is (are) not an example of a temperature inversion? Explain why not.

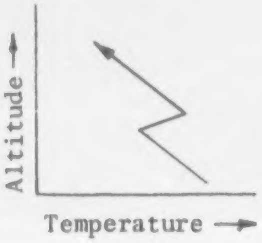
(A)



(B)



(C)



10. What is an inversion?

11. Under what circumstances will polluted air rise?

12. Explain, in terms of density, why polluted air becomes trapped by an inversion.